INTRODUCTORY
CHEMISTRY
AND PHYSICS

PICKARD AND RADOMSKY

test tubes

graduated cylinder
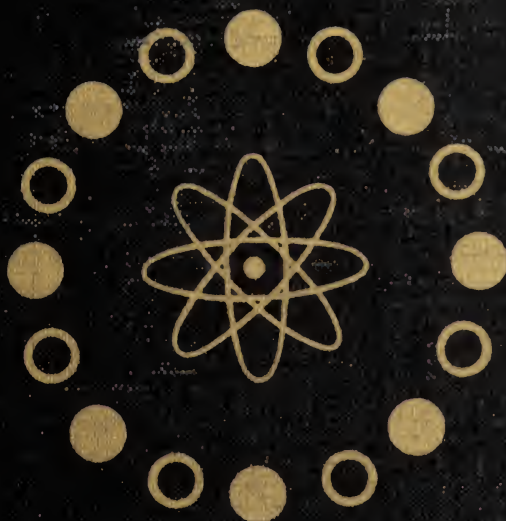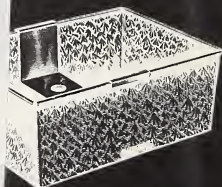
funnel

box of weights

pneumatic trough

wire gauze, asbestos centered

clay triangle

test tube rack

mortar and pestle

crucible and crucible cover

evaporating dish

Ex Libris
Universitatis
Albertaensis

QUAECUMQUE VERA

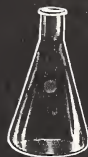calcium chloride tube

beaker

florence flask

erlenmeyer flask

gas bottle

liebig condenser

retort

blow pipe

burette clamp

forceps

tongs

deflagrating or combustion spoon

spatula

tube clamp
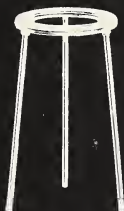
tripod

bunsen burner

test tube holder

INTRODUCTORY CHEMISTRY AND PHYSICS

INTRODUCTORY CHEMISTRY AND PHYSICS

NORMAN J. PICKARD
Assistant Principal, Crescent Heights High School, Calgary, Alberta

STEVE W. RADOMSKY
Head of Science Department, Strathcona Composite High School, Edmonton, Alberta

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Preface

PREFACE

This book is designed as a one-year course to introduce the subjects of chemistry and physics at the secondary-school level. The authors have presented the material, which is based on the modern concept of matter, in what they believe to be a logical arrangement. For best results, it is suggested that it be presented in the same sequence as it appears in the text.

Teachers and students alike should note that considerable emphasis has been placed on both practical laboratory work and the mathematics of science. It is recommended that a minimum of one class period in five should be spent in the laboratory. Some classes may have time for more laboratory periods.

It is hoped that the text will help students understand clearly the mathematical methods of solving problems. The steps involved in the solution of type problems have been explained in mathematical terms. At the end of each chapter there are exercises and problems which can be supplemented as time permits.

The authors are deeply indebted to Dr. S. G. Davis and Dr. D. B. Scott of the Chemistry and Physics departments respectively of the University of Alberta for their valuable contributions and critical review of this work.

Our sincere appreciation for guidance in this work is offered to Mr. J. Charyk, Dr. S. G. Davis, Mr. A. B. Evenson, Mr. O. Massing, and Dr. D. B. Scott as members of the Science Committee of the Department of Education of Alberta.

Grateful acknowledgment is also made to the teachers of science in the Senior High Schools of Calgary, Edmonton, and other Alberta centers for their helpful suggestions and the conducting of two years of pilot courses on which this work is based.

The authors also wish to express their gratitude to Miss F. M. Davis, Mrs. B. Detmold, Mrs. P. Mason, Mrs. S. W. Radomsky, and Mrs. M. L. Singleton who did much in the preparation of this manuscript for publication.

S. W. R.
N. J. P.
February, 1960
Calgary

THE ELEMENTS

ELEMENT	SYMBOL	ATOMIC NUMBER	ATOMIC WEIGHT	ELEMENT	SYMBOL	ATOMIC NUMBER	ATOMIC WEIGHT
Actinium	Ac	89	227	**Mercury**	Hg	80	200·61
Aluminum	Al	13	26·98	*Molybdenum*	Mo	42	95·95
Americium	Am	95	(243)	*Neodymium**	Nd	60	144·27
Antimony	Sb	51	121·76	**Neon**	Ne	10	20·183
Argon	Ar	18	39·944	*Neptunium*	Np	93	(237)
Arsenic	As	33	74·91	**Nickel**	Ni	28	58·71
Astatine	At	85	(211)	*Niobium*	Nb	41	92·91
Barium	Ba	56	137·36	**Nitrogen**	N	7	14·008
Berkelium	Bk	97	(245)	*Nobelium*	No	102	(253)
Beryllium	Be	4	9·013	*Osmium*	Os	76	190·2
Bismuth	Bi	83	209·00	**Oxygen**	O	8	16·0000
Boron	B	5	10·82	*Palladium*	Pd	46	106·7
Bromine	Br	35	79·916	**Phosphorus**	P	15	30·975
Cadmium	Cd	48	112·41	**Platinum**	Pt	78	195·09
Calcium	Ca	20	40·08	*Plutonium*	Pu	94	(242)
Californium	Cf	98	(248)	*Polonium*	Po	84	210
Carbon	C	6	12·011	**Potassium**	K	19	39·100
*Cerium**	Ce	58	140·13	*Praseodymium**	Pr	59	140·92
Cesium	Cs	55	132·91	*Promethium*	Pm	61	(145)
Chlorine	Cl	17	35·457	*Protactinium*	Pa	91	231
Chromium	Cr	24	52·01	**Radium**	Ra	88	226·05
Cobalt	Co	27	58·94	*Radon*	Rn	86	222
Copper	Cu	29	63·54	*Rhenium*	Re	75	186·22
Curium	Cm	96	(245)	*Rhodium*	Rh	45	102·91
*Dysprosium**	Dy	66	162·51	*Rubidium*	Rb	37	85·48
Einsteinium	E	99	(255)	*Ruthenium*	Ru	44	101·1
*Erbium**	Er	68	167·27	*Samarium**	Sm	62	150·35
*Europium**	Eu	63	152·0	*Scandium*	Sc	21	44·96
Fermium	Fm	100	(252)	*Selenium*	Se	34	78·96
Fluorine	F	9	19·00	**Silicon**	Si	14	28·09
Francium	Fr	87	(223)	**Silver**	Ag	47	107·880
*Gadolinium**	Gd	64	157·26	**Sodium**	Na	11	22·991
Gallium	Ga	31	69·72	*Strontium*	Sr	38	87·63
Germanium	Ge	32	72·60	**Sulphur**	S	16	32·066
Gold	Au	79	197·0	*Tantalum*	Ta	73	180·95
Hafnium	Hf	72	178·58	*Technetium*	Tc	43	(99)
Helium	He	2	4·003	*Tellurium*	Te	52	127·61
*Holmium**	Ho	67	164·94	*Terbium**	Tb	65	158·93
Hydrogen	H	1	1·0080	*Thallium*	Tl	81	204·39
Indium	In	49	114·82	*Thorium*	Th	90	232·05
Iodine	I	53	126·91	*Thulium**	Tm	69	168·94
Iridium	Ir	77	192·2	**Tin**	Sn	50	118·70
Iron	Fe	26	55·85	*Titanium*	Ti	22	47·90
Krypton	Kr	36	83·8	*Tungsten*	W	74	183·86
*Lanthanum**	La	57	138·92	**Uranium**	U	92	238·07
Lawrencium	Lw	103	(257)	*Vanadium*	V	23	50·95
Lead	Pb	82	207·21	*Xenon*	Xe	54	131·30
Lithium	Li	3	6·940	*Ytterbium**	Yb	70	173·04
*Lutecium**	Lu	71	174·99	*Yttrium*	Y	39	88·92
Magnesium	Mg	12	24·32	**Zinc**	Zn	30	65·38
Manganese	Mn	25	54·94	*Zirconium*	Zr	40	91·22
Mendelevium	Md	101	(256)				

*Rare Earths

Atomic weight in brackets denotes the mass number of the isotope of longest known half-life.

SECTION 1

An Introduction to Science

UNIT I - A SHORT HISTORY OF SCIENCE

"I am a part of all that I have met . . ." are the words attributed to the Greek hero Ulysses by the poet Tennyson. It is a true statement and applies to all of us no matter what our position in life. We are all part of all that we have met or seen—but if we follow the philosophy to its ultimate we are more than that.

As an illustration, you, if you live in a community served by electricity, do your studying in the evenings by electric light. Consider the question: How is this possible?

Your immediate answer may be that you can study by electric light because your father went to the store and bought some light bulbs. But that is only part of the answer. He could obtain these bulbs only because someone made a machine that would produce them according to the specifications of engineers who designed them. The engineers in turn had modified plans that originally came from Thomas Edison who gave us our first incandescent light bulb. But there would have been no light bulb if Michael Faraday, in the middle of the nineteenth century, had not developed a dynamo that gave us a source of continuous current produced by mechanical means.

In other words, you study by electric light, not just because your father paid the light bill, but because of all that has gone before in the development of our method of lighting.

The same holds true for all of the many advantages of our modern living. Today each of us is a part of all that we have seen; but more, we are a part of all those who have gone before and left their discoveries for us. We have a great debt to the past that can only be repaid by our contribution to the future.

← A stainless steel wire holds a gold-plated sphere that swings continuously as a pendulum in the main lobby of the General Assembly Building at the United Nations Headquarters in New York. This pendulum, as did the one in the original experiment of Jean Bernard Foucault, offers us visual proof of the rotation of the earth on its axis as the freely swinging pendulum shifts in its plane.

United Nations

CHAPTER 1 OUR DEBT TO THE PAST

In the years that you have lived, no doubt you have seen a section of a town or city change rapidly in its appearance. In fact, the change could have been so rapid that if you had been away from it for even a few years you would have failed to recognize the town on your return. These changes were the result of several causes. Possibly a few old buildings were remodeled. New houses were built on empty lots. Some buildings were demolished so that new structures could take their place. In some cases, new streets came into being carrying people along thoroughfares that had not existed a few months before.

Changes even more startling have taken place in the field of science in a relatively short space of time. As a result of theories that have been revised, new theories that have been introduced, and new avenues of study that have been explored, great changes have come about in the past few years. Possibly we do not realize or appreciate these changes because we accept the present without thought of the past.

It might help us to realize the extent and the rapidity of these new developments by considering the thought that if a person who has been dead for as short a time as one hundred years were able to come back to this life, he would have difficulty in recognizing the world in which he once lived. Indeed, it would not be the same world. During that time the automobile has replaced the horse-drawn vehicle. Trains and steamships have been remodeled. Planes have appeared in the sky. The telephone, radio, and television have become common means of communication. Miracle drugs now aid the physician.

Modern synthetic fabrics clothe the individual. The atom is yielding its secrets to provide us with a new source of power, and already man is talking seriously of visiting other planets. Truly the world has changed, and science and its application have made this change possible.

SCIENCE IS KNOWLEDGE

The word "science" comes to us from the Latin verb *scio* meaning "to know." There is also a Latin word *scientia* which means "knowledge." In both cases, in the Latin the idea of knowing includes much more than simple recognition. In fact *scientia* in Latin means knowledge in its widest sense. When man learned the reasons for the phenomena that he observed and when he understood the reasons sufficiently to make practical applications of the knowledge gained, then the world in which he lived changed. But all this took many years.

In the early portion of this course, while you are reading of the development of science, you will also be working with the projects at the end of each chapter. You should spend at least one day each week in the laboratory, which is the workshop of the scientist. These projects are designed to help you to become acquainted with the equipment that scientists use in the laboratory. You should also, through practice, develop some skill in using it. In the first two sections of this book, the laboratory work follows each chapter. In the third section the laboratory work follows each unit.

THE EARLY PERIOD

In the earliest of times, man's knowledge was meagre and its applications crude. As in any other period, however, important discoveries were made. Many of these, we believe, were the results of accident or trial-and-error methods. Long before there was any organized record of knowledge, man had learned the value of fire and how to ignite combustible material. Runners of a conveyance that today would be called a "stone boat" had been replaced by a crude wheel. Rawhide had been tanned. Coloring matter had been obtained from vegetables and berries, and early in his history man had learned some of the practical aspects of fermenting juices.

The development of knowledge was slow, for several reasons. Because of the lack of communication, each student of science had to depend on his own resources with no help from those outside his circle. For the same reason, much of the work accomplished died with the death of the individual or with the decline of the group with which he worked.

At first the explanation of natural phenomena observed was put on a mystical basis. While we realize today that this was useless to a world searching for knowledge, we must also realize that it was an advance over earlier groups that offered no explanation.

In some areas, though, where learning was well advanced, surprising strides had been made as early as 2000 B.C. From the findings of archaeologists, we know that the Babylonians were versed in mathematics and had instruments that enabled them to study the science of astronomy. In this field, they became sufficiently advanced that they could foretell the occurrence of a solar eclipse. Because of lack of communication, much of their knowledge died with the fall of Babylonia in 538 B.C.

THE PHILOSOPHERS OF GREECE

By 500 B.C. Thales of Miletus, one of the Greek philosophers who thought much about natural phenomena, concluded that the universe and all things in it followed natural laws. Therefore, once these laws were understood, the universe could be explained. While this may seem to us a very elementary statement of what we accept as a fact, it was at the time of Thales a great step forward in scientific thinking.

Democritus, another Greek philosopher, expressed the opinion that all matter was made up of tiny particles. On first observation, we might be tempted to say that this was similar to John Dalton's atomic theory of later years. We must remember, however, that Democritus was attempting to explain things as a whole while Dalton, after careful investigation and thought, was offering an explanation of definite substances. In the case of Democritus, like so many of the early philosophers, enthusiasm and speculation were far ahead of careful observation on a scientific level.

While scientific observation and work of a practical nature was not usually carried on by the Greek philosophers, Aristotle of Athens was a notable exception. About 400 B.C. Aristotle made one of the first classifications of animals that we have. True, the classification was limited to a few animals that the philosopher knew, but he made careful notes describing them and noted their differences. Possibly more important than the classification was his careful recording of his work. For many years following his efforts, these records served as a guide for those who followed the study of science.

One of the great scientists of this early period was Archimedes of Syracuse. True to the Greek philosophy, Archimedes, a mathematician, was much more interested in theory than in applied science. Some applications of his work were made during his lifetime, but his greatest contribution comes down to us today as the statement known as Archimedes' principle which deals with the buoyant force that liquids exert on bodies immersed in them.

THE EGYPTIAN SCIENTISTS

The Egyptians made a great contribution to the science of the day when they established the Museum of Alexandria in 250 B.C. This became much more than a collection of articles that had to do with historical events. Under the Ptolemies it was developed into a classical university as well as a school of practical research. Its library contained over 400,000 manuscripts. It had departments of literature, astronomy, mathematics, and medicine, and experimental work was carried on there. This was a sign of progress over the Greeks who glorified thinking but ignored practical applications. Although the Museum was destroyed by the Christians in 390 A.D., rebuilt, and again destroyed by the Mohammedans in 640 A.D., it was recognized as the center of learning for the known world for almost 1000 years.

An interesting aspect on some of its work can be seen in its study of alloys. The Egyptians were recognized as excellent workers in the metals such as gold, silver, and copper. In addition to these metals, there is evidence that research was carried on at Alexandria relating to the mixing of the less noble metals to produce cheap alloys that looked like silver and gold. Articles made from these would be within the economic reach of the masses. In the light of this it would seem apparent that costume jewelry, as it is known today, is certainly not new and that from the times of the Pharaohs until the present, in the words of Kipling, the Colonel's lady and Judy O'Grady have been sisters under the skin.

THE ROMAN CONTRIBUTION

With the fall of Greece the heritage of science fell to the Romans. Their main contribution to the advancement of science was an indirect one. To the Romans, pure knowledge of a scientific nature was of little value, but practical applications were useful. In an effort to sort out the useful from the useless in the light of their own values, much of the learning of the conquered nations was brought to Rome. Here for the first time in the history of civilization the work of east and west met that of north and south.

Fortunately for the cause of science, the works of Aristotle had been translated into Arabic, so that all was not lost when the barbaric hordes of the north overran the Roman Empire.

THE DARK AGES

With the fall of Rome the light of science in Europe burned low. Here and there a solitary scholar or a member of a monastic order still searched for truth.

It was during this time that a group who became known as the alchemists came into being in Europe. Actually their origin goes back to the Arabian school of science, but the group became better known for its efforts in Europe. Workers in the black arts, teachers of astrology, and casters of spells, they wielded great influence over a people already burdened by superstition. Their investigations were as varied as they were weird. By combinations of simple chemical reactions, incantations, and mystic rites they hoped to produce such wonders as the philosopher's stone, a stone that would change base metals into gold and prolong the life of all who possessed it. Like the purveyors of some patent medicines of later days, they searched for the Elixir of Life that would restore youth and cure all disease. When the confidence men of later years started selling "gold bricks," the game was actually an old one originated by the alchemists. In their enthusiasm they even hoped to find a universal solvent. Their lack of ability to reason scientifically is evidenced by the fact that in none of their writing is mention made of the securing of a suitable container for the solvent.

Fortunately, for the cause of science, there were a few of the alchemists that were sincere and honest in their attempts to discover truth. In spite of the lack of scientific methods as known today, certain substances were isolated. In addition, the more serious workers became dissatisfied with the

Fig. 1-1. An alchemist in his laboratory, grown weary of the hopeless task of changing "base" metals into gold. What pieces of equipment can you identify that are still used in the modern laboratory?

aims and methods of the group as a whole.

While there is no definite date that would mark a turning from alchemy, the year 1500 and the name Paracelsus should be noted. Paracelsus, a physician and the son of a physician, was one of the first to denounce the practice of alchemy. He noted that there was a relationship between cause and effect. He made a careful study of certain compounds found in nature and sought to establish their use in medicines.

His methods were far ahead of his time, and his life was filled with trouble from enemies who were followers of the old methods. Nevertheless, gradually some of the forgotten achievements were again remembered. The influences of the stars and incantations were discounted. The ability to compare things by standards of weight, length, and volume crept back into science, and the practice of alchemy began to decline.

THE MIDDLE PERIOD

The start of the middle period might be marked as the time man began to examine his methods of study.

Francis Bacon (1561–1626), one of England's greatest philosophers, was noted for his forceful presentation of the case for the inductive method of reasoning as applied to the field of science. Prior to the time of Bacon the work of the scientist was, for the most part, carried out in an unorganized fashion. Much of it was trial and error. This was noted by Bacon. His message to

the scientists of the day was to experiment, observe, collect all the known facts pertaining to the problem, and then to form a theory.

It would be impossible to point to any one individual and mark him as the "father of modern science." The names of two men, however, in fields that we would classify as physics and chemistry should be noted. They are important not only because of their discoveries but also because of their successful employment of the use of scientific methods of investigation. This did much to set the pattern for future progress.

One of the first men to do much to establish this new approach was Robert Boyle (1627–1691), son of the Earl of Cork, who devoted his entire life to scientific study. Strange as it may seem to us today, because this great scholar was close enough to the time of the alchemist, he still believed in the possibility of the transmutation of lead into gold. Nevertheless, because of his work in chemistry, elements were distinguished from compounds for the first time. In the field of physics he formulated the law that concerns the relationship of the volume of a gas to the pressure to which it is subjected. Boyle's methods were those which had been outlined by Bacon.

The second scientist that demands our attention lived during the eighteenth century. He was Antoine Lavoisier (1743–1794), a French chemist, who gave us our accepted explanation of combustion. With the careful use of the chemist's balance he demonstrated for the first time the quantitative determination of chemical reactions.

Lavoisier's methods, while based on the inductive process as suggested by Bacon, were expanded to include further application of the new theory. This grew into the scientific method of reasoning which we practice today in all of our investigations.

THE SCIENTIFIC METHOD

Any subject being investigated scientifically is dealt with in four distinct steps.

1. All the known facts, results of experiment and observation that concern the problem, are carefully noted.

2. From the known facts that have been observed, a reasonable explanation is formulated. This is an example of inductive reasoning, as advised by Bacon. In this case we pass from a particular fact to formulate a general theory.

3. From the general theory formulated as the result of the second step we may now reason to other particular cases. This is an example of deductive reasoning. In this way we may prophesy results of related problems.

4. Experiments must be devised at this point to prove or to disprove the original theory. The results of these experiments will determine whether the theory is acceptable as it stands, whether it must be modified, or whether it must be discarded.

Actually, in most cases, when step 4 is carried out it is found necessary to modify the theory. In practice this modification may, and usually does, continue for many years. The reason for this is that the scientist continues to study and experiment. The results of any one of these experiments may produce new evidence that will require changes in the statement of the theory.

THE LATER PERIOD

While man began to use the scientific method as a guide for his studies, not only did he add to his knowledge in the field of pure science, but he also began to make applications of his knowledge.

Less than one hundred years after James Watt discovered and harnessed the power of steam (1765), others had applied his principle to give the world its first steam railroads and steamships. Other industries also benefited, directly as well as indirectly, from this discovery and its applications. In a relatively short space of time the world was faced with the industrial revolution.

This example is but one illustration of the advance of science in the eighteenth and

nineteenth centuries. In that period discoveries were made, theories were evolved, and principles were established that changed the life of the world. When the Italian physician, Luigi Galvani (1737–1798), discovered that a current of electricity was possible, Alessandro Volta (1745–1827), another Italian scientist, was challenged to produce a current from his invention that we now call the electric cell. A few years later an English scientist, Michael Faraday (1791–1867), followed with a device called the dynamo which produced a current by using magnetic fields of force. In less than 130 years since Faraday's immortal experiment, electricity has become the servant of mankind.

It was also during this period that a Swedish chemist, Svante Arrhenius, developed his theory of ionization. While it was not accepted immediately by all the chemists of his day, it did provide answers to many of the questions that had been puzzling them. The theory was accepted, with some modifications, a few years later, and electrochemistry became an important study in the field of science.

Until this time men had believed the chemistry of nature was not for them to investigate. Friedrich Wöhler, however, a German chemist, was willing to face the opposition of his fellow-students and started serious study in this branch of natural science. In 1828 he successfully synthesized an organic compound called urea, and a new era in chemistry was opened. Because of his work we benefit from modern drugs, synthetic fabrics, plastics, organic dyes, and a host of other products which have become a part of our daily life.

With research being carried on continuously, particles smaller than the atom began to play important roles. As a result, our systems of communication have been completely changed. In addition, the breakdown of the atom has given the world an entirely new source of power. It has already been tried in war and is at present being tested in peacetime applications.

THE PROGRESS OF TODAY IS THE RESULT OF THE PAST

In your study of arithmetic you have learned that money makes money. This is especially true when the interest on the money loaned is computed on a compound rate. The scientific achievements of today are the results of compounding knowledge. In the early days, because of the lack of communication, the scientist was a lone figure working in his laboratory. In most instances he was a local student, unknown except within the radius of a few miles from his home.

With the advent of improved means of communication through the printing press, the railroad, and the telegraph, themselves products of scientific research, scientists who made discoveries worthy of note were sought by their fellow-scholars.

One characteristic of men who search for knowledge is their willingness to share their knowledge with their fellow-students. Unfortunately, during times of international unrest and war, this is not always possible because of the danger to the security of the nation. Nevertheless, it is interesting to note, with a few exceptions the story of scientific progress has been continuous and has rarely recognized international boundaries.

Because science has proved its own value, the scientist has been accepted. While superstition and prejudice still control the thinking of some people their influence has lessened to the point where it does not hamper true scientists. Governments recognize their work, and national research councils have been formed. In a short space of time private industry, too, has realized the value of their work, and today there are few large industrial concerns that do not use laboratories for the quality control of their products as well as for research.

In every instance the scientist is conversant with the work that has gone before him in his field. His debt and our debt to the past is indeed great.

THE VALUE OF STUDYING SCIENCE

Many high school students ask the question, "Why should I study science?" To the scientist the very fact that such a question is asked is considered as a good sign. The answer that is given here, it is hoped, will encourage students and give purpose to their studies.

For those who wish to follow a career in science the answer is self-evident. The men who will build the rockets and jets in the years to come are not the boys who rush out today and, after reading a magazine article, attempt to produce such an engine. The builders of tomorrow's scientific world are the students of today who study the basic principles already known and upon which future research depends.

There is also a fundamental value to the sincere student of the subject which is quite apart from the gaining of scientific knowledge about the world. If the subject and its methods are studied carefully, the student will acquire the habit of using the scientific approach as a guide to his thinking on any topic. If for no other reason, the acquisition of this habit alone makes the study of the subject well worth the time and effort expended.

THINGS TO REMEMBER

Science is knowledge.

The Greek philosophers made some contributions based on observation, but most of their work was unsupported theory.

The Museum of Alexandria was, for a thousand years around the time of Christ, the center of learning.

During the Dark Ages superstition and prejudice prevented any appreciable progress being made in scientific investigation. It was during this time that alchemy flourished.

The break from alchemy started about the latter part of the sixteenth century. Paracelsus, a Swiss physician, was one of the first to denounce the practice.

Most important sign of progress in science was the examination by man of his method of study. Bacon, Boyle, and Lavoisier were leaders in the development of the Scientific Method.

The Scientific Method is based on the following pattern:
1. Collect the facts concerning the problem by observation and experimentation.
2. Formulate a reasonable explanation of the observed reaction or phenomenon.
3. From the general theory prophesy the results of related reactions.
4. Test the general theory, modify if necessary, retest.

The application of knowledge gained opened new fields.

Our present state is dependent on the past.

The search for knowledge is an international one.

QUESTIONS

Part A

1: Give the derivation and meaning of the word "science."

2: State several reasons why progress in science was slow during the early period of its study.

3: Why do scientists today look back and say that even a false theory was a sign of progress?

4: What scientific discovery is credited to the Babylonians?

5: Examine a picture of any alchemist in his laboratory. Are there any pieces of equipment shown that are similar to equipment used today? Name or describe such pieces. What was the alchemist's source of heat?

6: While we do not use the methods of Paracelsus today, his contribution to science is still considered important. What was his contribution? Why is it still considered important?

7: State and discuss the steps followed in any scientific investigation.

8: What do we mean by "inductive reasoning"? Who first advocated that this method be applied to scientific investigation?

9: State the difference between inductive and deductive methods of reasoning.

10: Name the scientist that first recognized the difference between an element and a compound. State his definition of an element.

11: Why is the scientist accepted today?

Part B

12: Copper, silver, and gold were known to the ancients, while aluminum and iron which are more abundant and are more common today were not known. Explain the reason for this.

13: The alchemists made some discoveries that proved to be of value to the later scientists. We do not, however, consider the alchemists as true scientists. Why not?

14: Explain what is meant by the transmutation of elements. Give several reasons why the modern scientist was able to succeed in transmuting elements while the alchemist failed in his attempts.

15: Suggest several reasons why the alchemist gained such a hold over the people of his time.

16: It has been pointed out that the common man today enjoys more comforts of life than the nobles of the past. Justify this statement by naming things we consider essential today that were not even known as luxuries one hundred years ago.

17: List six ways in which science has contributed to *your* enjoyment of life.

18: List six ways in which science has added to your years of life or increased your security.

19: State instances where superstition or prejudice have hampered the cause of science, that is, the search for truth, in the ancient world. Suggest any instances where this has taken place in your own time or possibly in your own community.

20: Describe briefly some developments of modern science about which you have read from articles or advertising.

21: The scientific method of study should be of value in fields other than science. Suggest some problem that is of practical interest to you around your home or school, and outline a suitable plan for studying it, using the scientific method as a guide.

THINGS TO DO

From your social-studies material obtain an outline map of the world. Each time that a scientist is mentioned as coming from a certain country mark that country on the map by using a small figure. Below the map list the figures and follow them by the name of the scientist, the approximate date of his work, and his contribution. You will be amazed to find that science truly knows no borders and that the search for knowledge has been international.

IN THE LABORATORY

General Instructions

The laboratory is the workshop of the scientist. During your studies of Chapters 1, 2, and 3 you should become familiar with some of the apparatus and techniques that the chemist uses in his work. In later chapters you will apply some of this knowledge.

In addition to knowing the names and the uses for various pieces of equipment in your laboratory locker, there are certain fundamental rules that you should know and follow.

Neatness, and above all cleanliness, are of first importance if you are to be successful in your laboratory experiments. Glassware should be washed and rinsed in distilled water before it is put away.* The bench top should be sponged and dried before you leave the laboratory. A single drop of an acid left on the bench could prove costly to the next person using the laboratory by ruining clothing or causing painful burns.

You should listen to the instructions given to you by your teacher and you should follow them. You should be familiar with the directions in the text and never try experimenting on your own just to see what happens. Many accidents have resulted from such efforts.

It is advisable to check your apparatus, especially the glassware, each time you use it. A chipped test tube may lead to cut hands,

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* In the case of a shortage of distilled water, test tubes that have been washed in tap water should be allowed to dry in the air in an inverted position after as much water as possible has been shaken from them. Beakers may be treated in the same way, or dried with a clean paper or cloth towel.

while a cracked piece of glassware will likely break as soon as heat is applied to it.

Solids such as matches, litmus paper, filter paper, or other such objects, must *never* be thrown in the sink. Waste jars should be provided for the purpose of keeping sinks free from solids at all times.

Soluble substances and liquids, with the exception of concentrated acids and mercury, may be flushed down the sink. The water should be kept running during the process.

Be courteous and careful. Remember most accidents are caused by carelessness, and very often others beside yourself are involved in the accident. Your responsibility for a safe laboratory is not only to yourself but to those who work with you.

Your Apparatus

Most of the apparatus with which you will be working is represented by the endpaper diagram. Some pieces will be discussed in detail later; the rest are mentioned here.

The *beaker* is a glass vessel, usually pyrex, which is used for reactions involving liquids where a greater capacity than that of a test tube is required. Heat from an open flame should never be applied to a beaker. If the contents of a beaker are to be heated, the vessel should be placed on a section of wire gauze.

The *mortar and pestle* come to us from the early days of the alchemists. They are used to pulverize substances that have become caked or to mix two or more solid substances.

The *Florence flask*, usually made of pyrex glass, is used for many purposes in the

laboratory. In your studies you will find that we use it for holding reacting substances which must be heated and the gas given off collected. As in the case of the beaker, wire gauze should be used if the contents are to be heated.

The *deflagrating spoon* is a long-handled iron spoon used to introduce chemical reagents into jars filled with gas.

Tongs and *test-tube holders* are used to handle equipment that is hot.

The *test-tube brush* is an important piece of equipment that should be used freely in keeping equipment clean.

The *crucible and cover* are made of porcelain. The crucible is used to hold small amounts of solid substances that have to be heated to high temperatures. A flame may be applied directly to this piece of apparatus.

The *evaporating dish* is used in the recovery of a solid that is in solution by boiling away the liquid.

The *burette clamp* and *ring stand* are used to secure apparatus in position for the experiment. Be careful not to clamp glassware too tightly for if it is heated it will expand and possibly break.

Wide-mouth gas bottles are used to collect gases. The large mouth facilitates the introduction of reacting substances once the gas has been collected.

The *test tube* can be thought of as the trade mark of the chemist's laboratory. Usually made of pyrex glass, it is used as a vessel in which small amounts of solids or liquids may be heated. Small amounts of reacting substances may also be tested in the test tube.

The *Bunsen burner* is the source of heat that is used in most laboratories today. There are several types, but all of them mix air with a gaseous fuel which burns at the top of an upright barrel. By adjusting the gas and the air, various types of flames may be produced.

To light the burner, after it has been connected to the gas supply set the air adjustment to about half open and turn on the gas supply. Bring a lighted match up the side of the barrel to the top of the burner. If you hold the match directly over the burner it will probably be extinguished by the force of the gas.
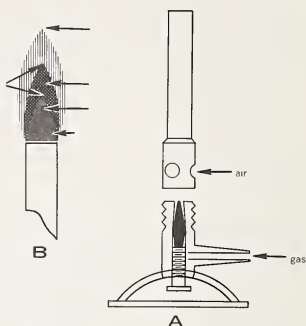

Fig. 1-2. The Bunsen burner and its flame.

Experiment with the burner and note the different types of flames that can be produced. By closing off the air completely you will see a yellow flame. This flame is inefficient and smoky. You can collect carbon from it by holding a porcelain dish in the flame for a few moments. It is not a very hot flame.

By opening the air control gradually you can change the yellow or luminous flame to one that is blue or nonluminous. You should be able to note an inner cone in this flame (see Fig. 1-2). The blue flame is the flame that is used in the laboratory as a source of heat. It is clean and indicates the complete combustion (burning) of all the gas delivered to the burner. By adjusting the gas flow, high and low flames may be produced.

ASSIGNMENT

Make a drawing of the burner and the nonluminous flame. Indicate the hot parts of the flame. Name the parts of the burner.

WORKING WITH GLASS

Because of the extensive use that is made of glass tubing in the laboratory, even the student of elementary chemistry should understand the basic techniques of handling it.
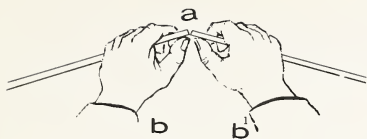
Tubing of suitable size is usually bought by the pound and comes in four-foot lengths. It may be cut by placing the tube on a flat surface and scratching it with a three-cornered file at its desired length. Hold the tube as shown in Fig. 1-3, with the scratch away from the manipulator and thumbs opposite the scratch. Bend the glass toward you using a quick motion while pulling apart at the same time, and the tube should break evenly.

The rough edges may be smoothed (fire-polished) by rotating the rough end for a few moments in a nonluminous Bunsen flame. Tubing may be bent to any desired angle by heating it in a Bunsen flame which has been flattened by means of a "wing-top" attachment.

After you have determined the point at which the bend is to be made, place the glass in the hot cone of the flame and turn it
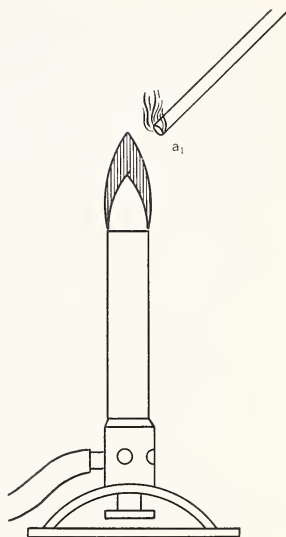
Fig. 1-3. The cutting and fire-polishing of glass. Hold thumbs firmly at *a* and pull the glass toward you at *b* and b_1. To fire-polish hold the rough end a_1 in the flame and rotate it until it has become smooth. Use soft glass for satisfactory results if you are working with a Bunsen burner.
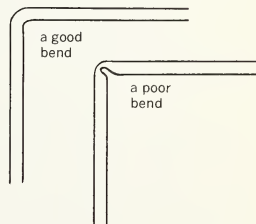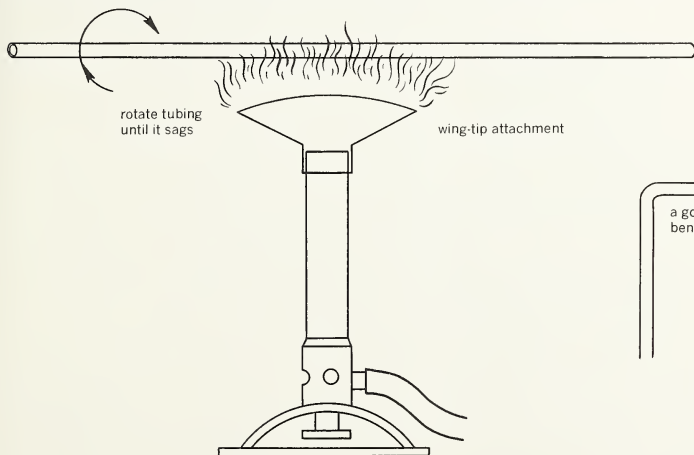

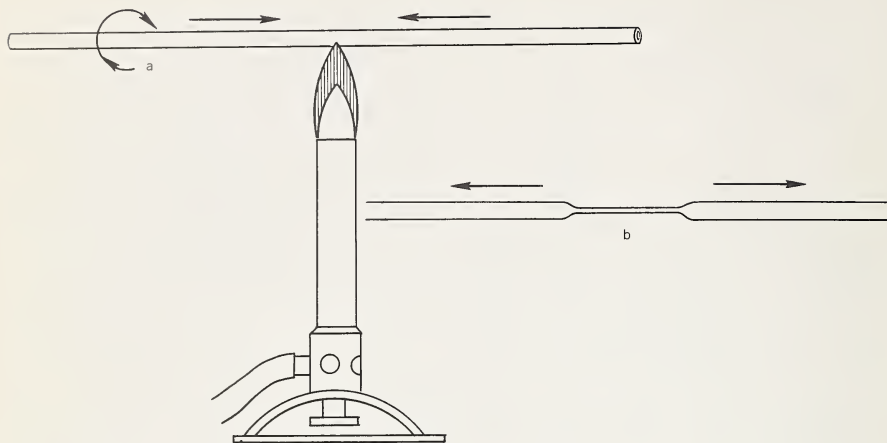
Fig. 1-4. Bending soft glass tubing.

Fig. 1-5. Making a glass jet. Rotate tube while heating as in *a*. Pull softened glass in opposite directions slowly and evenly as in *b*. When cool, cut glass to form two jets.

slowly, ensuring an even heat throughout the section. When the glass shows signs of sagging from its own weight, remove it from the flame and bend it to the desired angle. Place it on an asbestos mat and allow it to cool slowly.

If a jet is desired it may be produced by following the directions for the bending of tubing without the "wing-top" and, at the point that sagging occurs, the glass may be stretched by pulling the hands apart. It may help if the ends are pushed slowly toward one another before stretching the glass to allow the wall to thicken to about twice its original thickness.

Glass may be inserted into stoppers only when it has been fire-polished and is cool. The hole in the rubber stopper and the glass should be lubricated with water. The rubber stopper should be gripped in a towel between the thumb and forefinger of one hand and the tubing should also be wrapped in a towel before being gripped by the other hand. The tube will slip through the opening in the stopper if it is worked into the opening with

a rotary motion. Hold the tube only about one inch from the stopper, or you may break it.

Never put any pressure on the glass at the bend. Be careful to take precautions against cuts that may result from breakages.
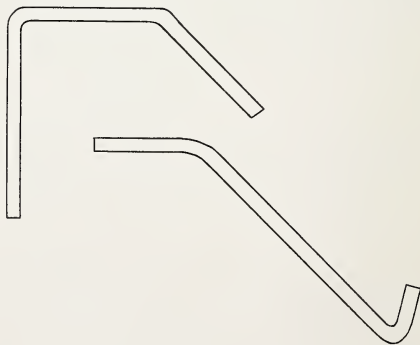

Fig. 1-6. Delivery tubes.

EXERCISES WITH GLASS

1. To make a stirring rod, cut a suitable length of solid glass rod (6 to 8 inches long). Fire-polish the ends until they are well rounded. Another type of stirring rod may be made by cutting a suitable length of glass tubing (6 to 8 inches) and, by the same method that you fire-polished the rod, treat the tubing until the end seals. Seal the other end.

2. To make a gas delivery tube, cut one piece of tubing 6 to 8 inches in length and, after fire-polishing the ends, shape according to the diagram. Cut a second length of tubing 10 to 12 inches long, fire-polish the ends and shape according to Fig. 1-6. After they have cooled set them aside for use later in the course.

    Note: "Soft" glass may be worked readily in a Bunsen flame. If pyrex glass is used, some difficulty may be encountered due to the higher temperature needed to soften it.

There are certain characteristics common to all scientists. Possibly one of the most outstanding is the possession of an enquiring mind. When the scientist observes a new phenomenon, when he is faced with a problem, or when he plans some application of a scientific principle, he is not satisfied until he has enquired into all the phases of the question.

Most people enjoy observing nature. Many will remark on the wonders they have seen and then go on to other things. But these people, who are so easily satisfied, are not the scientists who have given us the world of today. The scientist wants to know what is behind that which he has observed. As one who walks over a hill to see what lies beyond only to find other hills, the searcher after truth is led on and on and, like the explorers of the physical universe, when he can go no further he leaves his knowledge to those who follow.

It is not surprising, therefore, that very shortly after the scientist observed the earth on which he lived he started asking such questions as, "Of what is this earth composed?" "Can it be classified?" "What is the structure of the material?"

By the middle of the twentieth century, scientists believe they have the basic answers, although they are not foolish enough to think that they have the complete answers. They are willing to admit that many of their answers may have to be modified. Of this, however, they feel certain, that while their beliefs are expressed as theories those theories are based on observations which have led to successful applications of the knowledge gained.

In this unit we shall discuss some of the fundamental facts concerning matter—the material of which our world and all that is in it is made.

CHAPTER 2 MATTER: ITS CLASSIFICATION

At the beginning of any new course, the student may be somewhat alarmed as he looks ahead and sees the number of new words and ideas that he must learn. A course in science provides no exception. If the subject is to be understood, the student must not only add a great number of new words to his vocabulary but also, in the study of chemistry, master a system of shorthand that the chemist employs to write formulae and chemical equations. It is not the purpose of this introduction to frighten the student. It is, however, hoped that the student will realize that in order to master such a course, hours of serious study are required.

The word "matter" is a scientific term used by the scientist to describe anything that *has weight and occupies space.* From this definition one might think it a simple thing to recognize matter, and so it would be, if all matter could be seen. But we must remember that the definition applies to much that is invisible. Air, for example, is matter even though it cannot be seen. The same is true for some other gases. It does not differentiate between the visible and the invisible; if the substance has weight and occupies space it is considered matter.

STATES OF MATTER

Matter exists in three states. The states are solid, liquid, and gaseous. A solid is described as a substance that is rigid. It has definite volume and definite shape without being held by a supporting container. The liquid has no definite shape, but will flow. It takes the shape of the supporting container as far as its volume permits. Like the solid, the liquid has a definite volume under constant conditions. A gas has neither definite shape nor definite volume. It takes its shape from the vessel which contains it, and its tendency is to expand, filling any container in which it is placed.

Many substances, under different conditions, exist in all three states. Water, for example, as a liquid may be changed into ice, a solid, by lowering the temperature below the freezing point of water. Water may also be changed to steam, a gas, by raising the temperature above the boiling point of water. You will be able to think of other substances that exist in all three states.

When water, as well as other substances, changes from the liquid to the gaseous state, we frequently speak of the gas as a vapor, and we refer to the new state as water vapor, mercury vapor, or sodium vapor. Water may be vaporized by boiling, causing it to pass rapidly from the liquid to the vapor state. It may also pass to the vapor state by evaporation. A liquid that boils at a low temperature and evaporates readily is said to be a volatile liquid.

The boiling point of water may be defined as the temperature at which the vapor pressure of water equals the atmospheric pressure. This same holds true for the boiling of all liquids. The process of a vapor changing back to a liquid is called *condensation.*

The temperature at which the liquid becomes a solid is known as the *freezing point.* If we consider the reverse action, the temperature at which the solid becomes the liquid is known as the *melting point.* The freezing point and the melting point are

actually the same temperature, at which a state of dynamic equilibrium exists between a liquid and its solid state.

Because these terms, freezing, melting, boiling, and condensation, are used so frequently with reference to water, many students forget that they apply to other substances as well.

CLASSES OF MATTER

Chemistry is a special branch of science that is primarily interested in the composition of matter and how it reacts with other matter. In such a broad field there are necessarily many subdivisions in order that the subject may be organized. Matter itself consists of pure substances or a mixture of pure substances.

1. *A Pure Substance* – This is matter that is composed of only one substance. Each particle is like every other particle and therefore the characteristics of each particle will be the same as the characteristics of all the others. Pure salt, pure sugar, a piece of pure iron, and a sample of pure water are all homogeneous substances.

2. *A Mixture of Substances* – This is matter that is not the same throughout. The particles that compose the samples are not alike and therefore have different characteristics. Soil, concrete, and bread are examples of mixtures.

The component parts of a mixture can be in any proportion. For example, in a mixture of salt and sand we could have as much sand as we wished or as much salt as we wished. In some mixtures we are able to distinguish the component parts by their appearance, but in others we have difficulty in doing this. We cannot, therefore, consider the ability to be distinguished by the appearance of its component parts as one of the characteristics of a mixture.

Elements and Compounds – Pure substances have been further subdivided by the chemist into two large subclasses: elements and compounds. A detailed discussion of these will be found in a later chapter, so for the present we will *consider the element as a substance that has not yet been decomposed into a simpler substance by chemical means.*

The compound is a substance which is the product of the chemical union of two or more elements. The compound does not have any of the properties or characteristics of the elements of which it is composed.
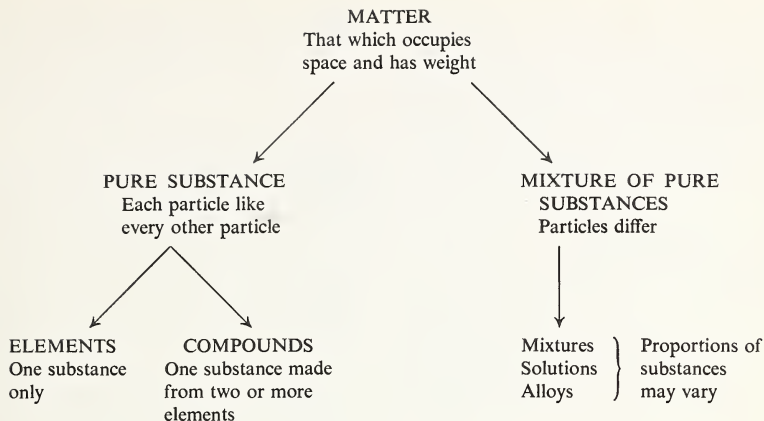
Solutions – These are a special type of a mixture of substances. Solutions, such as salt dissolved in water, while they may appear to be a pure substance, are mixtures of two or more elements or compounds. In reality they are special kinds of mixtures. Solutions are of great importance to the chemist because many of the reactions in which he is interested are carried out with one or more of the reacting substances in solution.

Solids, providing they are soluble, may be dissolved in liquids. Many solids are soluble in one liquid but not in another. The same holds true for liquids. Liquids which dissolve in each other are said to be miscible, that is, they form a solution that is the same throughout. Other liquids are immiscible with each other, for example, water and gasoline. Gases may also be dissolved in liquids or in other gases.

As well as liquid solutions, one can also have solid solutions. To obtain these the solids are heated until they melt, mixed together, and allowed to cool. The solids in such a case as this, of course, have to be miscible in each other. The solid solution is one type of alloy. Unlike pure compounds this example of an alloy may have a variable composition. As you advance in your study of chemistry you will find that there are other types of alloys.

The alloy is of great importance to industry. Very few of our finished metal products are made from the pure metallic element. Aluminum utensils and other aluminum products are made from alloys. Steel is an alloy of iron.

The chart on p. 21 will help to clarify the classification of matter.

MATTER
That which occupies
space and has weight

PURE SUBSTANCE
Each particle like
every other particle

MIXTURE OF PURE
SUBSTANCES
Particles differ

ELEMENTS
One substance
only

COMPOUNDS
One substance made
from two or more
elements

Mixtures
Solutions
Alloys
}
Proportions of
substances
may vary

PROPERTIES OF MATTER

Young people who are interested in automobiles are able to make a rapid identification of a car at a glance and name the car, its model, and its year. While most cars have four wheels and a body, subtle differences in their fenders, trim, or general lines make this identification possible.

Because the chemist is asked to identify many samples of matter, he must have a means of distinguishing between different substances, even if the substances appear to be identical. The characteristics of any substance, and it is the characteristics of the substance that make the identification possible, are known as its properties. No two substances have identical properties.

Physical Properties and Physical Changes – To refer back to the classification of the automobile mentioned above, we would agree that it was to be classified according to its appearance. A characteristic of the physical appearance of a substance which in no way changes the composition of that substance is called a physical property. The chemist makes use of many of these physical properties in his schemes of identification. The properties include such characteristics as state, color, hardness, density, taste, odor, malleability, ductility, solubility, melting point, and boiling point.

Changes that have to do with the form of the substance without changing its actual composition are said to be physical changes. Water changes to steam or to ice and the substance is still chemically the same; therefore the change is considered to be physical. A piece of wood 8 feet long is cut into two pieces, each 4 feet in length. The form is changed but the substance is still wood. The sawdust itself, while not considered wood by the carpenter, is still wood to the chemist. In a like manner, iron may be in the form of a ball, a bar, or even iron filings. As long as the sample has a density of 7.75 gm. per cc., with a melting point of 1530 degrees centigrade, it is an indication of iron to the chemist.

The crushing of rock, the melting of lead, the dissolving of sugar in coffee, and the evaporation of gasoline are all considered physical changes, for no new substance with a composition differing from the original has been formed.

Chemical Properties and Chemical Changes – Chemical properties have to do with the ability or lack of ability of one substance to react with another to form a third substance. When one substance reacts readily with another we say that the substances are active with respect to each other. If the reaction is slow we say the substances are inactive. If a substance will not react with any other substance, we say it is inert. In all chemical

changes, new substances are formed and there is an energy change, usually in the gain or loss of heat. Thus we would say that when a piece of wood burns, there is a chemical change, for the wood loses its identity as such and the new substances formed have properties quite unlike the original substance.

ANOTHER IMPORTANT CHARACTERISTIC OF A CHEMICAL REACTION

Chemists have found that when substances react chemically, the total weight of the reacting substances is the same as the total weight of the substance or substances formed.

EXPERIMENT

2-1. To illustrate that there is no change in the total weight of substances during a chemical reaction

Half fill an Erlenmeyer flask with a solution of salt and water. Pour 15 cc. of a solution of silver nitrate into a short test tube and place it in the flask in such a way that the two solutions do not mix (see Fig. 2-1). Stopper the flask tightly and weigh the flask and contents on a balance. Note the weight. Remove the flask from the balance and tilt it in such a manner that the two solutions mix. You will note that a chemical reaction takes place as the two clear solutions produce a white substance. This insoluble substance is called a precipitate. Place the flask on the balance again and you will note that there has been no change of weight.

By conducting hundreds of experiments in which chemical reactions were involved and finding that all the products of reaction were accounted for and there was no change in weight of the reacting substances and the products, the chemist then stated his conclusion in the law of conservation of matter. This law states: In any chemical reaction the total weight of the new substances formed is the same as the total weight of the substances that entered the reaction. From such a statement we can see that by chemical means matter can
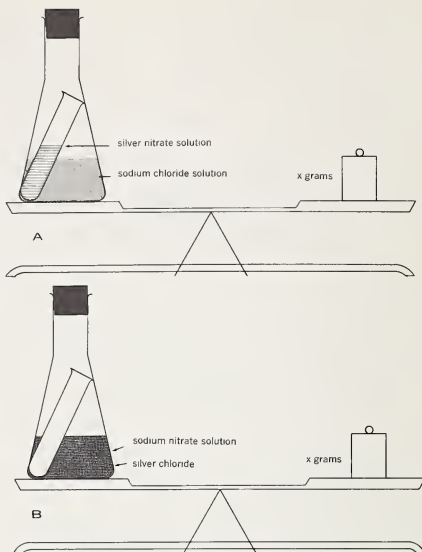

Fig. 2-1. Demonstration of the law of conservation of mass. A. Flask containing the silver nitrate and sodium chloride in separate compartments. B. After the flask has been tipped so that the two reagents have reacted. Note no change in the balance.

neither be created nor destroyed. It can merely be changed.

HEAT AND THE FORMATION OF CHEMICAL COMPOUNDS

When two substances react chemically to form a new substance there is a liberation or an absorption of heat energy. The reaction that liberates heat is called an *exothermic* reaction while the reaction that absorbs heat to force it to completion is said to be an *endothermic* reaction.

EXPERIMENT

2-2. To demonstrate an exothermic reaction

Half fill a test tube with a dilute solution of hydrochloric acid. Half fill a second test tube with a dilute solution of sodium hydroxide. Note the temperature of the contents of each test tube. Carefully add the contents of the first tube to the second. Note the temperature of the

resultant reaction. The temperature may be noted by the use of a thermometer or simply by touching the tube. This is an example of an exothermic reaction.

Some exothermic reactions require heat to start the reaction, but once the reaction has started the heat produced will carry it to completion. The burning of a piece of paper requires heat to raise the temperature of the paper to the kindling point or the point at which the paper burns. Once the paper is burning, the heat produced by the reaction is sufficient to keep the temperature well above the kindling point.

Endothermic reactions are difficult to demonstrate in the average laboratory because they usually take place at extremely high temperatures. Water, for example, when heated under pressure to an extremely high temperature breaks to its component parts, hydrogen and oxygen. Limestone, when heated to 800 degrees centigrade (800°C) and

Fig. 2-2. Demonstration of an exothermic reaction. *A.* A test tube of hydrochloric acid (dilute) at room temperature. *B.* A test tube of sodium hydroxide at room temperature. *C.* Note the temperature after the two reagents have been allowed to react.

held at that temperature for a few minutes, decomposes, forming calcium oxide (CaO) and carbon dioxide (CO_2). In the laboratory, this can be done by the use of a blow pipe or a blast lamp.

ENERGY

The heat given off or absorbed, as the case may be, in these reactions is a form of energy. *Energy is defined as the ability to do work.* There are many forms of energy. We are all familiar with electric energy, or the ability to do work through the utilization of electric power. Heat energy is utilized in the steam engine. One form of energy with which we may not be as familiar is *kinetic energy,* that is, the working power of a moving object. An automobile rolling down a hill with no power from the engine possesses kinetic energy. A bullet traveling through the air also has kinetic energy.

Many of our applications involve the use of several forms of energy. The automobile depends on the electric spark to produce heat to ignite the fuel that in turn produces *mechanical energy.*

Another form is *potential energy.* Gasoline, for example, as a liquid in a container is not doing any work and therefore does not come under the definition of energy. We know, however, that under the proper conditions it has great power. Because of the power it possesses we say it has potential energy.

MEASUREMENT, A TOOL OF THE SCIENTIST

So far we have been classifying matter according to its composition. This method, however, is not sufficient for the scientist. One does not advance very far in the study of physics or chemistry until not only the kind of matter but the question of how much matter enters into the discussion. This immediately raises the question of standards of comparison.

In the early period of man's history, rough standards of comparison were the only ones available. The length of the arm and the width of the hand were commonly used. While these would not be considered precise enough for our use they were considered satisfactory for the time. From these, and like units of measurement, developed many of the standards of comparison that are still used in the world today. Very few of the
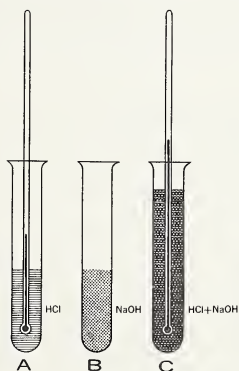
modifications of the original standards have any simple relationship to each other and, as a result, lengthy computation is required to convert one unit to its equivalent in a second unit. For example, consider the problem of figuring the number of inches in 1⅓ miles.

This problem was considered by the peoples of many countries, and by the early nineteenth century, French mathematicians developed what is now known as the *metric system of standards.* It is a system based on tens, as is our monetary system, and one in which the units of length, weight, and volume all have a simple relationship to each other. Within a very few years after its introduction, it was accepted for use in trade and commerce by all the civilized nations of the world except the English-speaking countries. It is accepted and used by the scientist all over the world today.

The unit of length is the meter. For practical purposes, the meter is considered as one ten-millionth of the distance between the north pole and the equator. There are, of course, subdivisions and multiples of the meter which are all based on tens. The scientist uses the centimeter (one one-hundredth of a meter) and the millimeter (one one-thousandth of a meter) more than he uses the kilometer which is one thousand meters.

The relationship between the unit of length and the unit of volume is established through the cubic centimeter (cc.). The basic unit of volume is the liter, a volume slightly larger than the American quart and smaller than the imperial quart. The liter contains 1000 cubic centimeters.* For this reason the cubic centimeter is also called a milliliter (ml.) which is one one-thousandth of a liter.

It is through this unit, the milliliter or cubic centimeter, that we establish the relationship of the unit of weight to the unit of volume. The unit of weight is the *gram,* which is the weight of 1 milliliter of water at a temperature

of 4°C and 760 mm. pressure (also expressed as 1 atmosphere of pressure).

The measurement of temperature that is used by all scientists is made by use of the thermometer graduated according to the centigrade scale. No doubt you are more familiar with the Fahrenheit thermometer where the freezing point of water is marked as 32 degrees above zero and the boiling point of water is marked as 212 degrees. The centigrade scale was chosen by the scientist because of its simplicity. Since water is the most important solvent that we have, the scale of the centigrade thermometer is based on its freezing point which is marked as zero and its boiling point which is marked as 100. The space between is measured into 100 equal divisions, each division being known as 1 degree centigrade. A comparison of the two scales is illustrated in Fig. 2-3.

There are times when the student is faced with the problem of converting figures from one of the systems of measurement to those of another system. This is usually done through the use of conversion tables. If the student knows certain basic relationships,
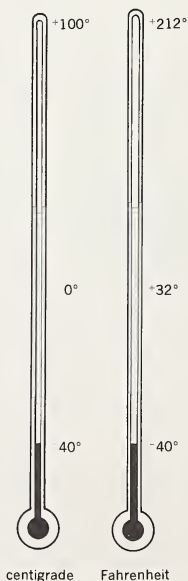

Fig. 2-3. Comparison of centigrade and Fahrenheit scales.

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* The liter is defined as the volume of pure water measured at 4°C and 760 mm. pressure which weighs 1 kilogram (kg.). In terms of cubic centimeters the liter is actually equal to 1000.027 cc. For practical purposes, however, we consider the liter as 1000 cc.
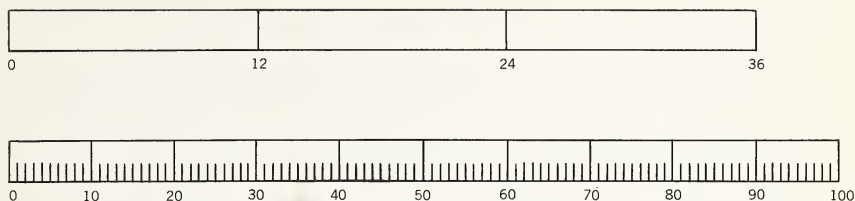
Fig. 2-4. Units of length compared in the English and metric systems.

however, he will always be able to work out the problem. A few of the more important relationships appear below and a more extensive set of tables is to be found in Appendix B page 230 of this book.

$$1 \text{ inch} = 2.54 \text{ centimeters (cm.)}$$
$$1 \text{ meter (m.)} = 39.37 \text{ inches}$$
$$1 \text{ kilogram (kg.)} = 2.2 \text{ lb.}$$
$$454 \text{ grams (gm.)} = 1 \text{ lb.}$$
$$1 \text{ liter } (l.) = 0.880 \text{ imperial quarts}$$
$$= 1.057 \text{ American quarts}$$
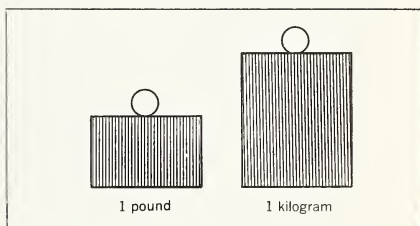
Fig. 2-5. There are 2.2 pounds in a kilogram.

THINGS TO REMEMBER

Matter has weight and occupies space.

Matter is found in three states: solid, liquid, and gaseous.

Matter is divided into two large groups: (1) pure substances and (2) mixtures of pure substances.

Pure substances may be subdivided into (1) elements and (2) compounds.

Mixtures of pure substances may be subdivided into three groups: (1) mixtures, (2) solutions, and (3) alloys.

Matter may be identified by its properties.

Physical properties concern matter as it exists.

Chemical properties are those properties of matter that have to do with its action or lack of action with other substances.

The law of the conservation of matter (sometimes called the law of the conservation of mass) states that in any chemical reaction the total weight of the reacting substances is equal to the total weight of the products of the reaction.

Heat is involved in chemical reactions. Exothermic reactions release heat during the reaction. Endothermic reactions absorb heat during the reaction.

Energy may be defined as the ability to do work. There are many forms of energy. These include heat energy, kinetic energy, electrical energy, chemical energy, and potential energy.

The metric system is the standard system of measurement accepted by the scientist.

QUESTIONS

Part A

1: Define matter. What particles of matter would you find it difficult to recognize?
2: Name the three states in which matter exists.
3: State the difference between a liquid and a solid, and the difference between a liquid and a gas.
4: What is meant by "a pure substance"? Name two subclassifications of matter that are pure substances.
5: Define an element.
6: Define a compound.
7: What is meant by a mixture of pure substances? Name three subclassifications of matter that fall into this class.
8: What do we mean when we speak of the physical properties of an element?
9: What do we mean when we speak of the chemical properties of an element?
10: State as many physical properties of iron as you can.
11: List eight properties that are useful in the identification of substances.
12: Which of the following reactions could be classified as physical changes and which as chemical changes? Give a reason for each of your choices.
    The dissolving of common salt in water.
    The melting of ice.
    The evaporation of water.
    The burning of paper.
    The freezing of iron.
    The cooking of meat.
    The production of light from a photoflash bulb.
    The melting of ice by the use of salt.
13: What is meant by saying that a reaction is exothermic?
14: Give an example of an exothermic reaction.
15: What is meant by saying that a reaction is endothermic?
16: Give an example of an endothermic reaction.
17: What do we mean by energy? Name and in a sentence describe four kinds of energy.
18: The scientist uses the metric system of measurement. What is the unit of length in this system? Name two subdivisions of this unit. What is the unit of weight? What is the unit of volume? What is the relationship between unit of volume and weight?
19: What is the great advantage of the metric system?

In the questions below involving computations give full statements to explain answer.

20: How many kilograms are there in 220 pounds?
21: How many feet are there in 100 centimeters?
22: How many grams are there in 12.25 kilograms?
23: How many cubic centimeters or milliliters are there in 11.55 liters?
24: How many liters are there in 6,475.35 cubic centimeters?
25: How many liters are there in 44 imperial quarts?

Part B

26: Describe how one type of alloy is made.
27: Explain the meaning of the word "miscible" giving examples of miscible liquids.
28: When you pound a nail it becomes hot. Is this an example of a physical or a chemical change? Explain your answer.
29: Following are five pairs of substances. In the case of each pair, state the properties of each substance that enable you to distinguish it from the other.
    a. wax and wood
    b. wood and iron
    c. copper and gold
    d. magnesium and silver (they both look alike when polished)
    e. water and gasoline
30: If you had a mixture of white sand and salt how could you separate the substances, recovering both in a dry condition?
31: Make a list of as many properties as you can that would enable you to distinguish between copper and sulphur.
32: What property would enable you to separate steel shot and lead shot?
33: How would you separate lead shot from tin shot if both kinds of pellets were the same diameter?
34: Some substances appear to resist change into states other than those in which they are normally used. Iron is such a substance. Where would you expect to find it in a liquid state and where in a gaseous state?
35. Give five examples of chemical changes.
36. In the demonstration of the law of the conservation of matter we stoppered the flask tightly. Apart from the fact that the stopper prevented the spilling of the liquids, suggest a reason why we did this.
37: If a piece of dry ice is left on the table it disappears completely. Does this refute the law of the conservation of matter? Give reasons for your answer.
38: What would be the easiest way to count out 10,000 pennies?

IN THE LABORATORY

Working with Liquids

Chemical graduates, commonly known as graduated cylinders or simply graduates, are commonly used to measure volumes of liquids.

The graduate is usually made of glass with the scale of measurement etched on the outside. You may also find it made of enamelware with the scale stamped on the inside. The scale is usually in milliliters (cubic centimeters) but the scale may vary depending on the capacity of the graduate.

Some graduates have double scales, one starting with zero at the bottom and indicating how much liquid has been poured into the cylinder, the second having the zero mark
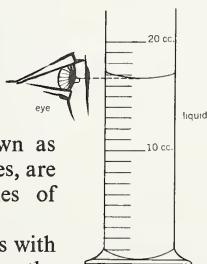

Fig. 2-6. When reading a volume of a concave meniscus (curved surface of a liquid such as water) the eye level should be on a horizontal line which would pass through the lower surface of the meniscus.

at the top and indicating how much liquid has been poured from the graduate.

If you pour some water into your graduate you will notice that the top of the water is not flat but it is curved. This curved surface is known as the *meniscus*. In the case of water and other liquids that wet the sides of the graduate the meniscus is concave.

When you read the volume looking through the glass sides be sure first, that the graduate is sitting on a level surface and second, that your eye is on the level with the lowest part of the meniscus.

Using Reagent Bottles – The chemist must be careful to see that he does not contaminate his reagents. For this reason, when he pours from reagent bottles he must never allow the glass stopper to come in contact with the desk or other surface. The one sure way to prevent this is to hold the stopper in the hand. This may be accomplished by removing the stopper as indicated in Fig. 2-7 and pouring the liquid without releasing the stopper from the fingers.
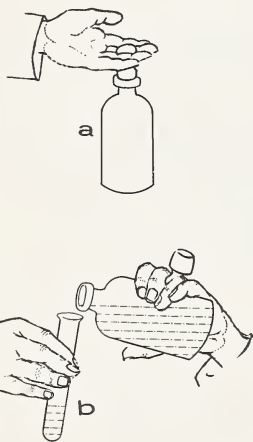
Fig. 2-7. To remove a glass stopper from a reagent bottle, grasp it as in *a*. Hold it between the fingers when pouring as in *b*. Make sure the lip of the bottle touches the lip of the receiving vessel so that no drops of liquid run down the side of the bottle.

When pouring, the neck of the bottle should be in contact with the rim of the test tube or beaker. The reason for this technique is that it prevents liquid from flowing down the outside of the bottle. If the liquid were acid or some other corrosive substance, painful burns to the hand would result.
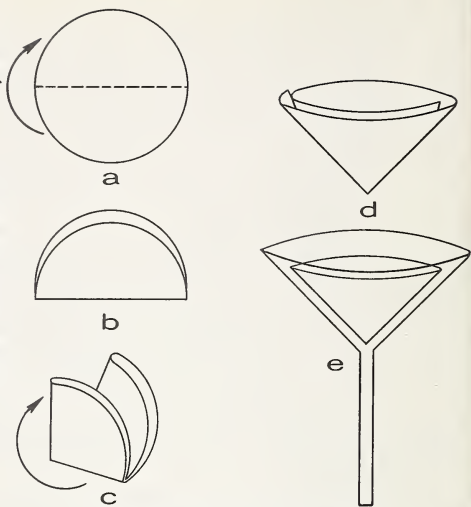

Fig. 2-8. Steps in folding a filter paper.

Diluting Sulphuric Acid – Make it a rule *always to pour concentrated sulphuric acid into water*. It should be added slowly while stirring the solution constantly. The action produces great quantities of heat, and if water were added to acid the contents of the vessel would likely be forced out by the steam formed.

Filtering – Insoluble material present in a liquid may be separated by the process known as filtering. Filter paper usually comes in circular shapes of various sizes. The paper is prepared for the funnel with which it is supported by folding it in quarters and then opening it out into a cone shape. It is then placed in the funnel and moistened with water to help it hold its position.

The liquid to be filtered is introduced into the funnel by means of a stirring rod (see Fig. 2-8), and as the liquid passes through the funnel, the solid material is held back by the filter paper. The solid material is now known as the residue and the liquid that passed through as the filtrate.

Heating Liquids in Test Tubes – When liquids are to be heated or boiled in test tubes the heat should never be applied to the
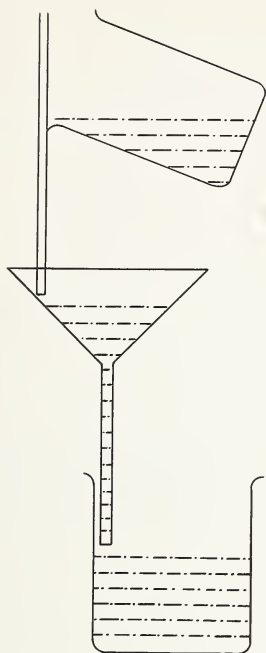
Fig. 2-9. Use a stirring rod to guide liquid into the funnel.
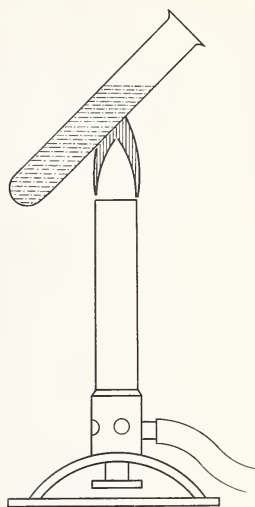

Fig. 2-10. Heating a liquid in a test tube. Agitate the test tube to keep the liquid in motion. Apply heat near the top level of the liquid.

bottom of the test tube. Instead, the test tube should be held at an angle and the heat applied just below the level of the liquid. To ensure even heating, the test tube should be agitated in the flame until a gentle boil or the desired temperature is reached. Be careful that the open end of the test tube is not pointed at yourself or your neighbor.

WORKING WITH SOLIDS

There are several solids with which you will be working this year that are quite corrosive to the skin (that is they appear to burn the skin). Never touch solids with the hands unless it is absolutely necessary and you are sure that it is safe to do so.

Powdered solids may be introduced into test tubes by means of a piece of stiff paper cut in the shape of a long narrow rectangle and folded to make a trough. Place the substance at one end and slide it to the end of the test tube where it may be shaken free.
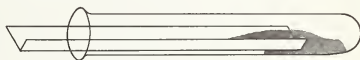

Fig. 2-11. Use a folded paper to place solids in a test tube.

EXERCISES WITH SOLIDS AND LIQUIDS

Take the amount of common salt that can be piled on a 25-cent piece and place it in a test tube. Pour 20 cc. of water, using a graduate to measure it, into the test tube and agitate the tube until the salt is dissolved.

Being careful not to let the stopper of the reagent bottle touch the bench, pour a few drops of a solution of silver nitrate into the test tube. What did you observe? Filter off the solid material. Have the teacher check your technique and results.
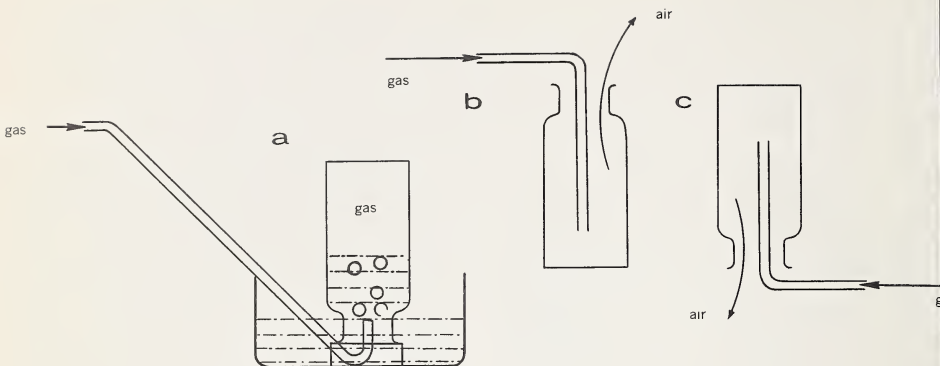
Fig. 2-12. The collection of gases. *A*. If the gas is insoluble or sparingly soluble in water it can be collected by the downward displacement of water. *B*. A gas which is heavier than air may be collected by the upward displacement of air. That is, the air is forced upward out of the receiving vessel. *C*. If a gas is lighter than air it can be collected by the downward displacement of air. That is, the gas when rising forces the air downward and out of the receiving vessel.

WORKING WITH GASES

Gases may be collected by one of three methods.

1. Gases that do not dissolve in water are collected by the downward displacement of water.

2. Gases that would dissolve in water but that are heavier than air we collect by the upward displacement of air.

3. Gases that are soluble in water but are lighter than air we collect by the downward displacement of air (see Fig. 2-12).

In the case of the first method you will now see why we made the delivery tube by bending the glass in Chapter 1. Take the longer of the two pieces of glass and follow these instructions.

Fill a pneumatic trough until the water level is about three-fourths of the way up the sides. Place a "beehive" support on the bottom of the trough. Fill a wide-mouth gas bottle with water and slide a glass plate over the top, making sure that no air bubbles remain in the bottle. Invert the bottle and lower it into the water. When the mouth is under the water, remove the glass plate and move the bottle on to the beehive (see Fig. 2-13).
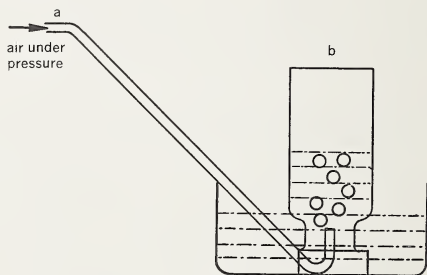

Fig. 2-13. Demonstration of the collection of a gas by downward displacement of water. Set up apparatus with bottle *b* filled with water. Insert delivery tube and blow air through the tube from point *a*. Note the water being forced out by entry of the gas.

Insert the glass tube under the beehive in such a fashion that the end of the tube points up into the bottle. Blow through the tube and note the results. When the water has been forced out of the bottle, raise the bottle from the beehive, keeping its mouth still under water, and slip the glass plate over the mouth. Turn the bottle right side up.

CHAPTER 3 MATTER: ITS COMPOSITION

How many times have you looked at an object that attracted your attention and aroused your interest and asked the question, "What is it made of?" That same question has been asked thousands of times in many different languages concerning the earth itself. Man has been searching for the building bricks of the universe for a long time. Because the scientist continued to search until he found the answer to many of the problems with which he was faced, we today have the modern world as our home.

Today we smile at some of the answers suggested by the early Greek philosophers, as the scientists were called in those days. Even an elementary-school student knows, for example, that water is composed of two substances, hydrogen and oxygen. The Greeks, however, suggested that the world itself was made up of four elements of which water was one and fire, earth, and air were the other three. Because of lack of careful observation and experimentation, men accepted the idea suggested by the Greeks. The acceptance is seen in our literary works where a storm, for example, is described as the raging of the elements. So well established was this idea that it was not until the seventeenth century that scientists established the fact that the Greeks were mistaken and science came closer to the answer it sought.

One of the chief difficulties in securing an answer to the question was lack of equipment and knowledge necessary to break down the substances into their component parts. Gradually, though, scientists progressed toward their goal. By 1700 12 elements, that are still considered elements today, were isolated.

Robert Boyle, mentioned earlier, is considered to be the first scientist to give us the basic definition of the element. In the year 1661 he wrote that the elements were the practical limits of chemical analysis and were substances incapable of decomposition by any means with which scientists were acquainted at that time.

Today the element is considered a pure substance that has not been decomposed into a simpler substance by chemical means. The phrase "by chemical means," while not included in Boyle's definition, is an important one. In 1910 Madame Curie, a French scientist, discovered an element which we call radium. This element, through a process called radiation, disintegrates naturally into simpler substances. Over a period of years a quantity of it turns into helium and lead, two other elements.*

Since the time of Madame Curie's discovery, other scientists have been able to break atoms into simpler substances by a device called the cyclotron. Such changes, however, are not chemical reactions, but enter into a field of nuclear science, a branch of the subject distinct from chemistry. Therefore the phrase "by chemical means" is important in our definition of an element.

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* On p. 56 you will be told that helium belongs to a family of elements that are inert, that is, they do not form compounds by uniting chemically with other elements. Yet here we see that radium breaks down to produce two elements, lead and helium. On the surface this may appear to be a contradiction. Both statements, however, are correct because radium is not a compound of lead and helium but is an element. Radium through the process of radiation, which is not a chemical change, disintegrates into two different elements, helium and lead.
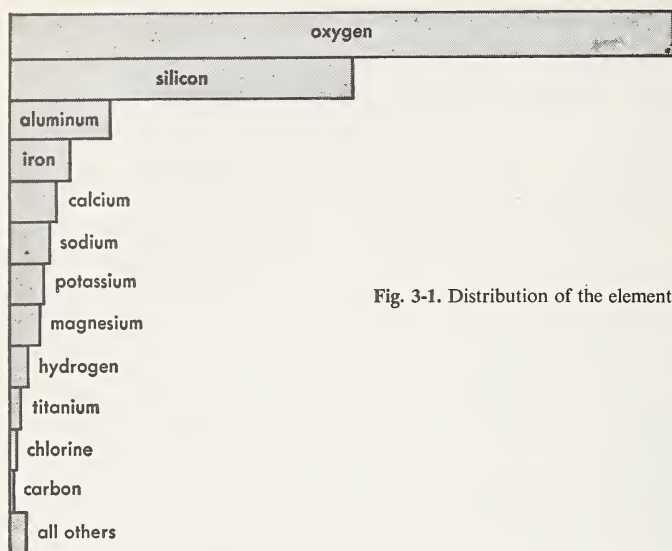
Fig. 3-1. Distribution of the elements in nature.

Today we have isolated and identified 89 elements that are found in nature and have synthesized 13 others.*

In this age, when we are conscious of the part played by science in our lives, we are all familiar with the names of some of the more common elements such as copper, sulphur, mercury, bromine, hydrogen, and oxygen. Possibly we are not quite so familiar with their classifications.

Elements are divided into two large groups known as metals and nonmetals. While we shall not consider their basic differences at present, each of the members in these groups has certain physical as well as chemical characteristics that cause it to be placed in one of the two categories. Metals, for example, have a lustrous appearance. They are good conductors of heat and electricity. They are malleable and ductile, that is, they may be pounded into thin sheets and drawn into fine wire. Nonmetals, on the other hand, with one or two exceptions do not have luster and do not conduct heat and electricity. They cannot be pounded into thin sheets nor drawn into fine wire.

Most of our elements, at ordinary temperatures, exist in the solid form. Two elements, mercury a metal and bromine a nonmetal, are liquids at ordinary temperatures. Several common elements including hydrogen, oxygen, nitrogen, fluorine, chlorine, neon, and argon are common elements which exist in the gaseous form at ordinary temperatures.

THE DISTRIBUTION OF THE ELEMENTS

We find that the earth's crust, including the atmosphere, contains various percentages of each of the natural elements known to man. Fig. 3-1 will give you some indication of the wide differences in the abundance of the elements. Do not think that this chart represents the elements in the free or elemental state. Most of them are in chemical union with other elements in the form we call compounds.

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* Scientists continue to produce new elements. It is also possible that some of the elements made synthetically at present may be even yet discovered in nature. Reports are now appearing that traces have been found in nature of plutonium 94.

COMPOUNDS

While elements are the basic building blocks of nature, they are not usually found in the free state as elements. They usually exist in chemical combination with other elements in a form known to the chemist as a compound. We may, therefore, define *a compound as a substance composed of two or more elements chemically united.* A little later we shall see that the proportion by weight of each element is always the same for each sample of the compound formed.

When a compound forms, a chemical reaction takes place, as the resulting substance does not possess the properties of any of the substances forming the new compound. For example, chlorine, a nonmetal, and sodium, a metal, are both extremely poisonous elements. By a simple chemical process these two elements may be united to form a compound called sodium chloride, which is common table salt, and instead of being poisonous is a necessary part of our diet.

EXPERIMENT

3-2. To demonstrate the formation of a compound

Take a sample of iron filings and a sample of sulphur. Note that the iron filings are attracted to a magnet. Note that the sulphur, a yellow substance, will dissolve in a liquid called carbon disulphide.

From your stock of these substances take $1\frac{3}{4}$ grams of the iron filings and 1 gram of sulphur and place them in a crucible as indicated in Fig. 3-2. Mix the two substances thoroughly. Heat the crucible and its contents strongly in the flame of a Bunsen burner.

In a few moments you will see that the mixture of the two elements will glow brightly and continue to glow even after the flame has been removed. After the glow has subsided and the crucible cooled, grind the contents in a mortar and examine the substance. You will see that a new substance has been formed. It is not soluble in carbon disulphide and it is not attracted by a
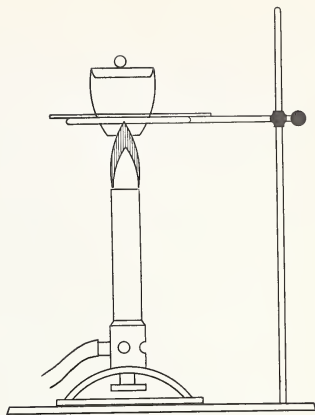

Fig. 3-2. Heating iron and sulphur in a crucible to produce iron sulphide. Cover the crucible and apply the heat using the hottest section of the flame.

magnet. In other words it has a new set of properties unrelated to those of the original substances. From this we can conclude that a chemical reaction has taken place and a compound formed.

LAW OF DEFINITE PROPORTIONS

If we carried out the above experiment a number of times making exact notation of the weight of iron used as compared to the weight of sulphur used, we would find that in each case the new substance produced, called iron sulphide, contained the same proportions of iron and sulphur by weight.

Chemists working with other substances have found that this is always true, and it has led to the formation of the statement known as the *law of definite proportions* which reads, *all samples of any given compound contain the same elements in the same fixed proportion by weight.*

This understanding is one of the foundation stones on which chemistry is built. It means that the chemist may be sure that a compound is a definite substance with properties of its own. The properties of that compound will be the same in his laboratory as in a laboratory on the other side of the earth provided the conditions are the same.

THE ATOM

Once man learned that elements and compounds were two different classes of matter, he turned his attention toward another problem that had been considered by the early Greek philosophers, the basic structure of matter.

The Greeks, as we have read, suggested a theory that all matter was composed of tiny particles. They even named these particles "atoms," meaning indivisible. Their reasoning involved the taking of a piece of matter, gold for example, and dividing it and subdividing it until the final speck was so small that it could no longer be subdivided and still be gold. This final particle was to the Greeks an atom.

Because man was interested in what he considered the more practical aspects of science, the whole question of the structure of matter was nearly forgotten until the last of the eighteenth century. By 1808 John Dalton, an English schoolmaster who was interested in the natural sciences, published a theory concerning the structure of matter that proved to be the basis for man's progress in the years that followed. While we do not accept the entire theory that Dalton produced, it did at the time give the scientist a definite hypothesis concerning matter.

In brief, Dalton's theory claimed:

1. The smallest particle of an element is known as an atom.
2. Atoms of any single element are alike in size, shape, and weight.
3. Atoms are indestructible as such and cannot be changed in any way.
4. Compounds are formed by the uniting of atoms in simple ratios.

Today we know that the atom can be smashed and that all atoms of the same elements are not exactly alike even if they have the same chemical properties. However, Dalton's theory led to many of the discoveries of later years, and upon it we have based our present definition of the atom as *the smallest particle of an element that can take part in a chemical reaction.*

ATOMIC WEIGHTS

One of the most important contributions Dalton made to the study of science was his postulate that all atoms of the same element had the same weight. It follows, of course, that the atoms of each element have a different atomic weight from the atoms of any other element.

Like Lavoisier, who ten years earlier demonstrated the value of quantitative chemistry, Dalton realized the necessity of establishing definite values to represent the weights of the various atoms. After a short time he realized that this was an impossible task for him, but by the analysis of certain compounds he was able to assign a series of relative weights for the atoms involved. Dalton was not noted for his ability as an experimenter, and many of his values were later proved to be in error. His attempt, however, gave the chemist a starting point for one of science's greatest achievements of the nineteenth century.

During his experiments Dalton realized that hydrogen appeared to be the lightest of the elements. He assigned the weight of 1 to its atom. By calculation he concluded the weight of oxygen would be 15.88. Later chemists, realizing that oxygen would be much more useful as a reference element than hydrogen, decided to base the comparison not on the weight of hydrogen but on that of oxygen and assigned the weight of 16.0000 to oxygen. This gave the weight of 1.0080 to hydrogen.

In a like manner, by analysis of many compounds, weights of other elements were established and checked. As a result of the modification of Dalton's idea the chemist has the atomic weight, as compared with oxygen, for each element with which he works. These weights may be found on the periodic chart of the atoms inside the back cover of this book. You will notice that the weight is the figure placed below the symbol O for oxygen, and the number 22.997 appears below the symbol Na for sodium. For our calculations we use the nearest whole number to represent the weight of the atom. Thus hydrogen becomes 1, sodium

becomes 23, potassium 39. There are two exceptions usually made to this general rule. Because of the fact that the weight of chlorine is so close to the middle point we usually consider its weight as 35.5 in calculations.

You will notice that no unit such as grams, kilograms, ounces, or pounds is assigned to the weight of the atom. The given weights are pure numbers and express a relationship only. This enables the chemist to use the unit of his choice when calculating theoretical yields of chemical reactions in the laboratory or in industry.

By definition, therefore, we can say *the atomic weight of an element is the weight of 1 atom of that element compared to the weight of 1 atom of oxygen taken as 16.0000.*

The value of using small whole numbers for relative weights may be readily realized when we consider the actual weights of the elements and the difficulties that would be involved in using them in computations. Since the time of Dalton, scientists have been able to compute actual weights of the atoms. For example, the actual weight in grams of an atom of hydrogen is 1.67×10^{-24} or

0.000,000,000,000,000,000,000,001,670

grams. The weight of an atom of oxygen is 2.7×10^{-23} or

0.000,000,000,000,000,000,000,027

grams. By the use of relative weights the chemist can avoid the cumbersome numbers illustrated above. The student, however, must always realize that the assigned atomic weights are purely relative and not let himself think that the weight of 1 atom of hydrogen is 1 gram or 1 ounce, but that it simply weighs 1/16 of an equal volume of oxygen under the same conditions.

SYMBOLS

From as far back as the fifteenth century, scientists have been using symbols to represent elements or what they believed to be elements. At first, each scientist or group of scientists had their own set of symbols (Fig. 3-3). You can see that some of the

Fig. 3-3. Symbols used by the ancients were cumbersome and varied from one locality to another.

symbols were quite complicated. As the study of science became more universal, it was realized that much confusion could be avoided if this method of representation were standardized and simplified. In the year 1813 a Swedish chemist, Jöns Jakob Berzelius, began using letters to represent the elements, and from his idea has grown the system of symbols that is used in the world of science today.

The system is based upon the representation of the elements by the use of one or two letters from the English or the Latin name of the element. It is necessary in some cases to use the second letter of the English word, and in others the Latin word, in order to distinguish elements that have names starting with the same letter. For example, consider the three elements: carbon, cobalt, and copper. All start with the letter C. Therefore we may use the letter C to represent carbon. Cobalt will be represented by the letters Co, but as copper also begins with the letters Co we take its Latin name cuprum and its symbol becomes Cu. In all cases the single letter symbol is a capital, while in elements represented by two letters, the first letter is capitalized and followed by a small letter.

In each case, the symbol represents 1 atom of the element and it also represents 1 atomic weight of the element. Thus, if we have the symbol O representing oxygen, we have 1 atom of oxygen with the atomic weight of 16. Other examples would be 1 atom of sulphur, S, with the atomic weight of 32, 1 atom of aluminum, Al, with the atomic weight of 27.

The elements are not all of equal importance. Many of our elements, while they have been identified, have little practical value at present. Out of the 89 natural elements you will hear about 40 in this course and you should know the symbols for these. They are included in the table at the front of this book and should be memorized.

In the fourth column the atomic weights are included. These become known to the chemist through use and are not memorized.

MOLECULES

The molecule of a compound is the result of the chemical union of two or more atoms. Hydrogen and chlorine unite to form the compound, hydrogen chloride. This may be represented by the chemist as $H + Cl \rightarrow$ HCl. It shows that 1 atom of hydrogen reacts chemically with 1 atom of chlorine forming the new substance, hydrogen chloride, that does not have any of the properties of hydrogen or chlorine but a set of properties of its own.

The molecule is, therefore, defined as the smallest particle of a substance that can exist by itself and still retain the chemical properties of that substance. In the case of the hydrogen chloride mentioned above, if we subdivided the substance into smaller and smaller particles we would eventually come to the time when we had 1 atom of hydrogen in union with 1 atom of chlorine. That would be a molecule of hydrogen chloride.

Some compounds have hundreds of atoms making up the molecule. Many molecules have more than one atom of a single element in the molecule. Water is an example of such with 2 atoms of hydrogen to 1 atom of oxygen. Dalton suggested that in such a case as this we indicate it by writing a small subscript following the element that has more than one atom in the compound; therefore, the molecule of water will be written as H_2O.

Chemists have also discovered that certain elements cannot exist as single atoms. There are seven of these common enough to demand our attention at this point. They are oxygen, hydrogen, nitrogen, fluorine, chlorine, as well as bromine and iodine when in a gaseous state. In each case we find that they have 2 atoms to the molecule, and the molecule is therefore represented with a subscript $_2$ following the symbol. The formula for the molecule is therefore O_2, H_2, and so on.

In all other cases, under normal conditions the symbol of the element is used to represent a molecule.

THINGS TO REMEMBER

The element is defined as a pure substance that cannot be decomposed by any chemical means known at the present time.

Robert Boyle, in the middle of the seventeenth century, was the first to recognize the difference between an element and a compound.

Today we know of 89 natural elements and 13 synthetic elements.

Compounds are the result of the union of two or more elements.

Each compound has properties of its own differing from those of the component elements.

The law of definite proportions. All samples of any compound contain the same elements in the same fixed proportion by weight.

The atom is the smallest particle of matter that can take part in a chemical reaction.

Each atom is represented by its symbol and each symbol represents 1 atomic weight of that atom.

A molecule is the smallest part of an element or compound that possesses the properties of that element or compound and may therefore be identified as such.

The active element gases: oxygen, nitrogen, hydrogen, fluorine, chlorine, and bromine and iodine in the vapor form have diatomic molecules and their molecular formula is O_2, N_2, H_2, F_2, Cl_2, and Br_2 and I_2. Other elements in their natural state are considered monatomic.

It is possible to have an atom or a molecule of an element but it is not possible to have an atom of a compound.

QUESTIONS

Part A

1: How does the definition of an element that we accept today differ from the definition given to us by Robert Boyle?
2: Why do we include the phrase "by chemical means" in the definition of the element today?
3: In what century did man first recognize differences between elements and compounds?
4: How can radium be considered an element if it decomposes into helium and lead with the passage of time?
5: Name a metal that is a liquid at room temperature. Name a nonmetal that is a gas, one that is a solid, and one that is a liquid at room temperature.
6: If aluminum is such an abundant element as shown by Fig. 3-1 on page 32 why wasn't it used a hundred years ago?
7: A compound and a mixture both contain more than one substance, yet only the compound is considered a pure substance. Explain.
8: What are the properties of a compound in relation to the properties of the elements that make up the compound?
9: State an important characteristic of all compounds in relation to their composition by weight.
10: State the law of definite proportions.
11: Describe the atom as we know it today.
12: What does the symbol of the atom represent in addition to an identification of the substance.
13: Why was hydrogen chosen at first as a basis for atomic weights?
14: Why do we use oxygen as our basis of comparison today?
15: Who was responsible for starting to standardize chemical symbols?
16: How do standardized symbols help the cause of science?

Part B

17: What is a molecule? Give the symbols for a molecule of oxygen and a molecule of water.
18: Give the symbols for an atom of potassium and an atom of oxygen.

19: Is there such a thing as an atom of water? Explain your answer.
20: Discuss Dalton's beliefs concerning the size of the atom, man's ability to divide it, and his theory of atoms of the same element. State the modern views on the same subjects. Why is Dalton remembered in the history of chemistry?
21: We know that the weight of the atom is so small it means nothing to us. Yet we use the term gram molecular weight of a substance. Explain the statement: The gram molecular weight of hydrogen is 2 grams and the gram molecular weight of oxygen is 32 grams.

IN THE LABORATORY

Working with the Chemical Balance

A chemical balance is indispensable to anyone who is attempting a serious study of experimental chemistry. Because of its high degree of precision it will weigh objects within its range with great accuracy. Because it is a precision instrument it should be handled carefully.

There are many types of balances varying in price from under $100 to over $11,000. Fig. 3-4 is a simple representation showing the basic parts of most of these balances.

The central support of the balance is a post (PO). The metal beam (AA_1) is balanced on top of a wedge-shaped knife edge (K_1) that

may be raised or lowered by a rod running through the post. There is also a knife edge at each end of the beam (K_2 and K_3). The pans (P and P_1) are suspended below the knife edges (K_2 and K_3). A control knob (C) may be turned so that there is no weight on any of the knife edges, and when weights or objects are being put on the pans or when the balance is not in use it should be left in this position.

After the object to be weighed and the weights are placed on the pans, the control knob should be turned to put the balance in a weighing position and the indicator needle
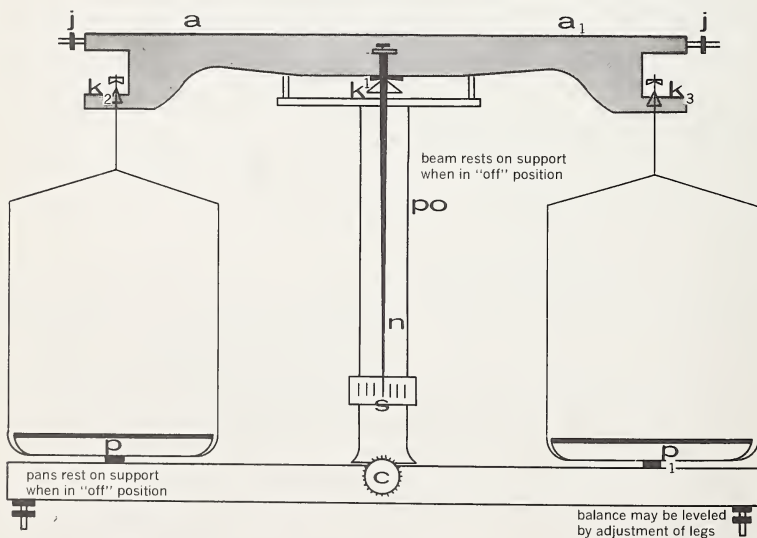

Fig. 3-4. Basic parts of a chemical balance.

(N) should be observed. If the needle swings to the left of the scale (S) it indicates that there is too much weight on the right-hand pan. If the needle swings slowly an equal number of divisions to each side, weight and object are in balance. The weights should then be added and the results noted.

If a balance is to remain in good condition and if accurate results are to be obtained, there are several rules that must be observed.

1. Remove any dust with a soft hair brush.
2. Objects to be weighed must never be placed directly on the pan. Solids should be placed on a watch glass that has been weighed previously. When the total weight of the glass plus that of the substance has been obtained, the weight of the watch glass is subtracted to give the net weight of the substance. Liquids are placed in a small beaker that has been weighed previously and later the weight of the beaker will be subtracted to obtain a final result.
3. Objects to be weighed should be free of contamination.
4. Objects to be weighed should be cool. This will prevent air currents from affecting the final weight.
5. Objects and heavier weights should be placed in the center of the scale pan.
6. The weights should be handled by tweezers as moisture from hands could affect the outcome.
7. Never place weights or objects on the pan while the balance is in the weighing position, that is, on its knife edges.
8. Always return weights to the box when finished weighing and make sure that the balance is in the "off" position.

The chemist usually uses weights that come to us from the metric system. The average box of weights contains units of 100 grams, 50 grams, two weights of 20 grams, a 10-gram weight, a 5-gram weight, two 2-gram weights, and a 1-gram weight. These are usually made of brass or nickel. In addition, there is usually a series of aluminum weights graded in milligrams, one one-thousandth of a gram. They range from 500 milligrams (.5 grams) to 10 milligrams (0.01 grams).

AN EXERCISE IN WEIGHING A SMALL OBJECT

With the balance in the "on" position check that it is truly balanced before putting anything on either pan. This is done by observing if the swing of the needle is the same on both sides of the scale. If it is not, ask your teacher to adjust the balance by turning the adjusting nuts (J) until the swing is equal.

With the balance in the "off" position place the object on a previously weighed watch glass on the left-hand pan. Estimate its weight and place the weights from the box on the right-hand pan. Raise the beam to the weighing position and observe the needle.

Add or subtract weights as necessary, making sure the balance is in the "off" position before the weights are changed. Use tweezers to add or remove weights.

When the balance is complete, record the weights and total them.

Calculate the net weight of the substance by subtracting the weight of the watch glass.

Have the teacher check the results while the object and weights are still on the balance.

Turn the balance to the "off" position and return the weights to the box.

CHAPTER 4 MATTER: ITS STRUCTURE

As we read a textbook on the history of any subject, discoveries seem to fall into place in a neat pattern. We must not be misled, however, by the printed account of the progress of science. In most cases, years of experimental work, observation, and study were required before the next piece appeared to take its place in the over-all design. Between the time that Dalton suggested the atomic theory and the introduction of the Atomic Age, statements of theories have been modified continually to fit the new facts discovered. In some instances, theories have had to be discarded completely even though they represented years of work. Such are the discouragements that the scientist faces in his search for knowledge. It is reported that Albert Einstein, one of the most notable scientists of the twentieth century, when asked to comment on a new discovery simply remarked, "The new facts have smashed my old ideas like a hammerblow." But instead of trying to fight an observed fact he used it to correct his own thinking in building a new theory to fit the situation. Such men have given us our modern world.

At this point in our study of science, physics and chemistry lose their identity in the pursuit of the question, "What is the structure of matter?" For purposes of this course, the answer centers around three particles which we find in the atom. As the particles themselves are electrical in nature we shall look back briefly at the growth of the theory that led to our conception of the atom as a unit of electrical charges.

From the time that Thales, the Greek philosopher, rubbed the substance we know as amber with a piece of cloth and found it would attract small bits of dry material, to the school student of today who rubs a hard rubber pocket comb and notes that it attracts small pieces of paper, people have been interested in electricity.

To demonstrate the effect of electric charges, rub an ebonite (hard rubber) rod with a piece of wool cloth. Move the end of the rod close to a light object (such as a pith ball) suspended on a silk thread. You will see the ball attracted to the rod. If allowed to touch the rod the ball swings away from it.

If a glass rod is rubbed with silk and brought close to the pith ball the ball is now attracted to the glass rod (see Fig. 4-1).
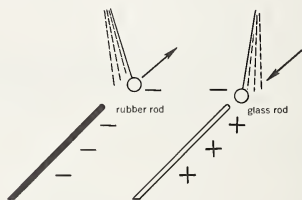

Fig. 4-1. Once the pith ball touches the rubber rod it is repelled. It will be attracted, however, by the positively charged glass rod.

Scientists of the middle nineteenth century came to the conclusion that there were two kinds of electrical charges, *negative and positive*. They observed that **unlike charges attract while like charges repel.**

Later they concluded that *a substance was electrically charged due to the presence or the absence of negatively charged particles. The negatively charged particles are called electrons.*

When you rub the ebonite rod with wool, electrons collect on the rod. When the rod touches the pith ball the electrons are transferred to the ball, and both ball and rod are negatively charged. Therefore they repel each other.

When you rub the glass rod with a silk cloth, electrons are removed from the glass thus leaving it positively charged. Therefore it will attract the negatively charged pith ball.

However, such an explanation came only after scientists discovered the nature of the electron.

Toward the end of the nineteenth century Sir William Crookes, an English scientist, passed a current of high voltage electricity through a glass tube from which all but a small portion of the air had been removed.

Crookes noticed that a beam of light appeared to be traveling from the negative terminal (the cathode) to the positive terminal (the anode). The beam could be seen more clearly when a piece of metal coated with a substance that fluoresces in the presence of an electric current was introduced into the tube.

The remarkable thing Crookes noted, however, was the effect of a magnet on the beam that traveled from the cathode (negative terminal). Depending upon the position of the magnet, the beam could be deflected from its path being attracted or repelled by the magnet.

EXPERIMENT

4-2. An experiment that led to the discovery of the electron

Connect a Crookes tube to a source of high voltage electricity (such as an induction coil energized by a 6-volt battery). Note the passage of the beam of light. Place a magnet in positions as indicated by Fig. 4-2.

Crookes was unable to explain this strange phenomenon of a "beam of light" being deflected by a magnet, and passed on the information to another eminent English scientist, J. J. Thomson, who was head of the

Cavendish Laboratory of Cambridge University, for explanation.

Thomson realized the magnitude of the problem but he also realized that the answer would no doubt provide a great step forward in the study of matter. Rays of light were just not deflected by magnets. This beam must be something more than a ray of light.

After twenty years of research Thomson gave the world his answer. The answer refuted the belief held up till that time that the atom was the smallest particle of matter. Thomson claimed the beam of light was composed of tiny particles of negative electricity which he called *electrons.* Furthermore, he discovered that electrons were present in the atoms of all elements and that all electrons were alike.

This discovery opened the field for further study, and in 1911 a student of Sir J. J. Thomson's, Sir Ernest Rutherford, conducted experiments that showed the presence of positive charges also present in the atom. These charges are due to the presence of *protons.* The proton is an actual particle of matter which is positively (+) charged. In 1923 the American scientist, Robert Millikan, for many years director of the California Institute of Technology in Pasadena, was awarded the Nobel prize in physics. His research showed the weight of the electron to be 1/1837 the weight of the proton.

In the year 1932 a third component of the atom was discovered by an English scientist, Sir James Chadwick. This unit he named the *neutron.* It may be described as a single unit that is electrically neutral. The neutron is approximately the same size and weight as the proton. Both are considered to be millions of times smaller than the whole atom and therefore they have a very high density.
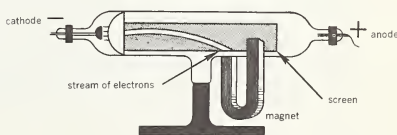

Fig. 4-2. Crookes tube.

With the successful conclusion of Chadwick's work, science had a new picture of the atom. That picture is still used today and describes all matter as being composed of electrons, protons, and neutrons.

Since the time of Chadwick other components of the atom have been discovered, In our study of matter, however, we will concern ourselves only with the electron, the proton, and the neutron.

INTRODUCTION TO THE PERIODIC CHART OF THE ELEMENTS

Since Aristotle first made the classification of the fauna of his day, scientists have found that the organizing and classifying of knowledge aids them in their thinking. It is quite possible that if the student who studies this course gains nothing more from it than that, and later practices it, the course will have been justified.

As men found new elements, several attempts were made to classify them according to various standards. Of all the early attempts the one that appears most significant to us today was the classification of Dimitri Mendeleeff (1834–1907) whose arrangement of the elements according to their atomic weights led to *the periodic law*. This law was based on the fact that when the elements are arranged in order of their atomic weights, we find elements with similar properties appearing at regular periods or intervals. Mendeleeff arranged his table so that those elements that had similar properties appeared in a vertical group on the table.

At the time the table was made Mendeleeff realized there were some imperfections in it. In order to keep together the "families," those elements that had similar properties, it was necessary to make a few exceptions to his rule of ascending atomic weights. Cobalt had to be placed before nickel, and tellurium had to be placed before iodine. After the discovery of the inert gas, argon, it was found that it had to be placed before potassium.

Mendeleeff, as well as other scientists, felt that while such a table was basically correct because of the recurring properties, the order of the elements did not depend on atomic weights.

The man who provided the answer to the problem was an English scientist, H. G. J. Moseley, who in 1913 worked on the structure of the atom as had Rutherford. By using the X-ray spectra Moseley demonstrated that each succeeding atom carried one additional proton in the nucleus. To the total of the protons in the nucleus the name of Atomic Number has been given.

The atom is electrically neutral because the number of electrons is equal to the number of protons.

If you look at the chart of the atoms on the inside back cover of this text you will notice that the elements progress in order of their atomic numbers. This means that each succeeding element differs from the one that precedes it by one electron on the outer or planetary rings and, therefore, also by one proton in the nucleus. There are other differences, but these are the ones that we are interested in at present.

This leads us to a new *definition of the element as being a substance composed of atoms all having the same atomic number*.

THE ARRANGEMENT OF PARTICLES IN THE ATOM

Neils Bohr, the Danish physicist and Nobel prize winner, and Sir Ernest Rutherford gave us the first picture of the structure of the atom in 1913. They compared it to a miniature solar system made up of a central nucleus consisting of protons and surrounded by electrons traveling in orbits at a relatively great distance from the nucleus.*

Actually, this means that most of the atom itself is space. You will have difficulty realizing this if you try to picture the atom as an infinitesimal particle. To aid us in this

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* Since Bohr presented his original theory the ideas of atomic structure have been modified considerably. We still think of the electrons occupying the allowed energy levels suggested by Bohr, but the positions of the electrons with respect to the nucleus is now much more uncertain. However, the simple Bohr theory is adequate for our purposes.

conception of the true relationship existing in the atom, the artist has combined his efforts with the scientist to give us the picture of an atom enlarged so that the nucleus measures 500 yards in diameter. In a case such as this, the diameter of the outer orbits of the electrons would be 3000 miles. Each electron, of course, would be only a few yards in diameter.

The number of protons and neutrons in the nucleus as well as the number of electrons and orbits depends upon the element.

When Bohr first pictured the atom the neutron was unknown. By assigning the weight of 1 to the proton, there was a dis-

crepancy between the calculated weight of the atom and the actual atomic weight. This difference was resolved once Chadwick discovered the neutron. Bohr's theory was modified to fit the new picture and in this way he gave us the diagram we use today (Fig. 4-3).

It has been found much simpler to represent the atom on the plane surface of a piece of paper rather than three-dimensionally as given us by the artist. While using the plane surface, however, we must remember that it is not, to the best of our belief, a true representation. The simplest atom we have is that of hydrogen.
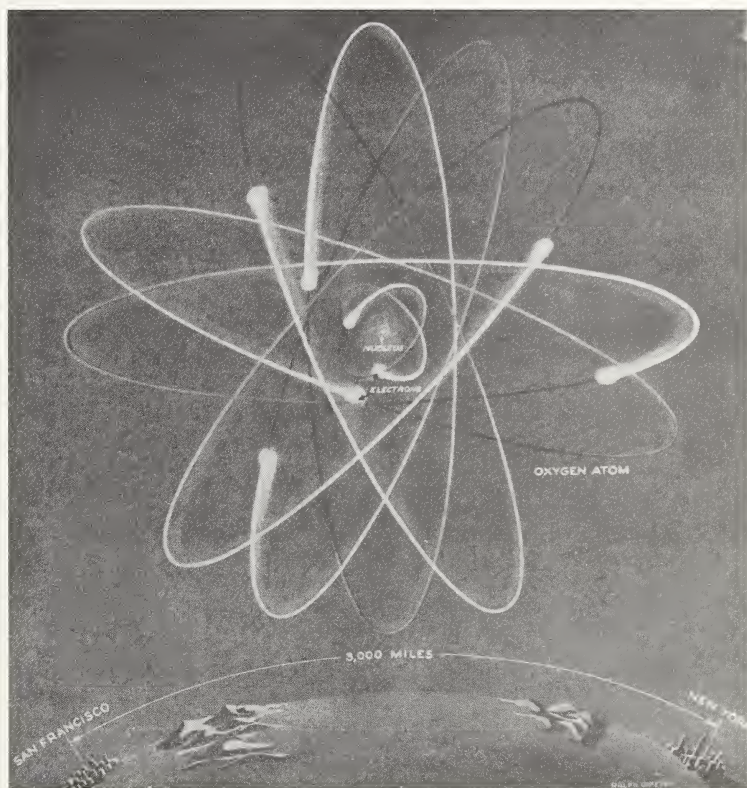

Fig. 4-3. If the nucleus of the oxygen atom were 1500 feet in diameter, the diameter of the path of the outer electrons would be 3000 miles. What does this indicate to you?
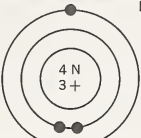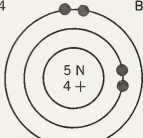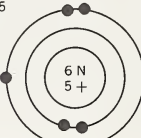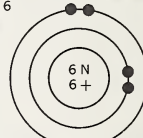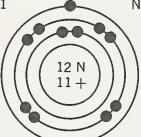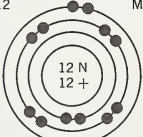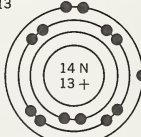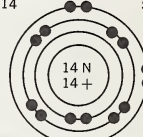
	Group 1	Group 2	Group 3	Group 4	
Period 1	1 H 1 + Hydrogen 1.008				
Period 2	3 Li 4 N 3 + Lithium 6.94	4 Be 5 N 4 + Beryllium 9.02	5 B 6 N 5 + Boron 10.82	6 C 6 N 6 + Carbon 12.01	
Period 3	11 Na 12 N 11 + Sodium 22.997	12 Mg 12 N 12 + Magnesium 24.32	13 Al 14 N 13 + Aluminum 26.97	14 Si 14 N 14 + Silicon 28.06	

Fig. 4-4. The first eighteen elements by atomic number. Numbers are shown

Hydrogen – The nucleus of the atom of hydrogen consists of a single proton and, therefore, as all atoms are electrically neutral there will be but one planetary electron. It can be represented according to Fig. 4-5.

Fig. 4-5.
The hydrogen atom.
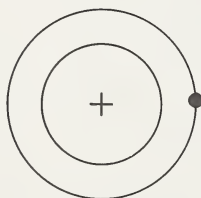
You will note that a circle is placed around the nucleus and the next circle represents the orbit of the planetary electron.* This orbit is referred to as a ring or as a shell.

Helium – Element number 2 is helium, which may be represented as having a nucleus of two neutrons and two protons. Like all other atoms it is neutral electrically; therefore there will be two electrons on the planetary ring. The atom may be represented as in Fig. 4-7.†
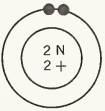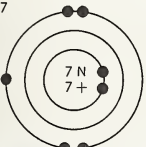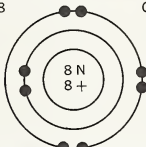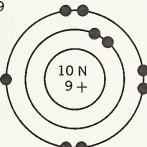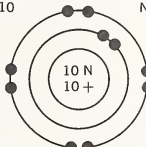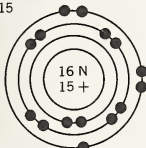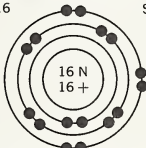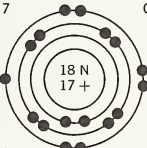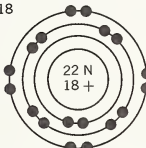
A GUIDE FOR THE CONSTRUCTION OF DIAGRAMS OF ATOMS

Rather than attempting to memorize the structure of each atom, you should learn the meaning of the atomic number and the atomic weight and their use in preparing the diagrammatic representation of the atom itself. Note how they apply in the case of helium the diagram of which we already have (Fig. 4-7).

A   B   C   D
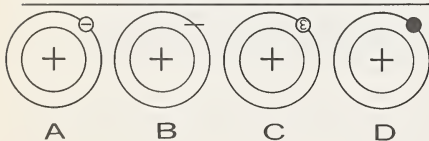
* Planetary electrons are represented by various symbols in different texts. They all, however, indicate the same thing, a negative charge. The symbol chosen for this text is the heavy black dot (Fig. 4-6, example D).
† In this text the proton in the nucleus is represented by a + sign; the neutron is represented by the letter N.

Group 5	Group 6	Group 7	Group 0
			2 He 2 N 2 + Helium 4.003
7 N 7 N 7 + Nitrogen 14.008	8 O 8 N 8 + Oxygen 16.000	9 F 10 N 9 + Fluorine 19.00	10 Ne 10 N 10 + Neon 20.18
15 P 16 N 15 + Phosphorus 30.98	16 S 16 N 16 + Sulphur 32.06	17 Cl 18 N 17 + Chlorine 35.457	18 A 22 N 18 + Argon 39.944

in the upper left corner, weights in the lower right corner of each section.

Atomic weight of helium = 4
Atomic number of helium = 2
Atomic number gives us:
1. Number of protons in the nucleus.
2. Number of electrons in the planetary rings.

The difference (atomic weight less the atomic number) = 2

This figure gives the number of neutrons in the nucleus.

As each neutron represents the weight of 1 and each proton represents the weight of 1, the total of these must be the atomic weight. In our representation of these atoms we always round off the atomic weight to the nearest whole number. You may ask how does the decimal fit in to the atomic weight if protons and neutrons are said to have the weight of 1. The answer to that question does not lie within the scope of this course but, if you would care to do a little research on your own, you will find information under the heading of *isotopes* in any advanced chemistry text.

But now we must add something to our diagram. While the scientist has still much to learn concerning the arrangement of the protons and the neutrons in the nucleus, he does know something concerning the arrangement of electrons in the planetary rings. To the chemist, especially, this is of great importance as we shall see in the next chapter.

All evidence seems to point to the fact that each planetary ring will only hold a certain maximum number of electrons. In this course we are concerned only with the first 18 elements shown on your atomic chart as periods 1, 2, and 3.

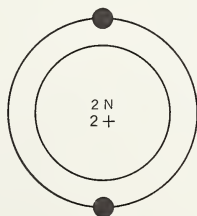The elements in period 1, hydrogen and helium, have only one planetary ring. This

Fig. 4-7.
The helium atom.

2 N
2 +

ring holds a maximum of 2 electrons. It is usually referred to as the K ring.

The elements of period 2, lithium through neon, have 2 planetary rings. The first ring, already mentioned as the K ring, holds its 2 electrons while the second ring known as the L ring holds a maximum of 8 electrons.

The elements of period 3, sodium through argon, have 3 planetary rings: the K ring with its 2 electrons, the L ring with its 8 electrons, and the third ring known as the M ring with a maximum of 8 electrons in period 3.

When we go on to period 4, potassium through krypton, another ring is added. Here, however, we shall find that the M ring takes on additional electrons up to a total of 18, but that study will be taken in a more advanced course.

Lithium – Period 2 starts with the element lithium, atomic number 3, atomic weight 7. This, of course, means 3 protons in the nucleus, 3 electrons (2 on the K ring, 1 on the L ring) and $7 - 3 = 4$ neutrons in the nucleus.

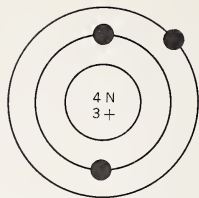You will notice that the element neon, number 10, has the full quota of electrons on

Fig. 4-8. The lithium atom.

its L ring and therefore sodium, number 11, will have to introduce one electron on the M ring.

For a complete picture of the diagrams of the elements 1 to 18 see Fig. 4-4.

Now that you realize the value of these numbers, you can see that our modern version of the periodic chart of the elements contains this information for each of the elements included. Some of the synthetic elements do not have atomic weights because the scientist can, in their manufacture, give them any weight he desires within certain limits. But for the elements which you will be studying in your high-school chemistry you will notice that each block in the atomic chart contains several pieces of information valuable to you.

The ring structure

Number of electrons in K ring	2		Atomic number = number of protons in nucleus
		11	= number of electrons in planetary rings
Number of electrons in L ring	8		
			Symbol represents one atomic weight of the element
Number of electrons in M ring	1	Na	
		23	Atomic weight compared to the weight of oxygen as 16.000. We usually go to the nearest whole number in calculations, always in diagram representation of the atom.

THINGS TO REMEMBER

Sir William Crookes, experimenting with high-voltage current, passed it through a glass tube from which most of the air had been evacuated. He noticed what appeared to be a ray of light coming from the cathode. This ray could be deflected by a magnet.

Sir J. J. Thomson later explained the phenomenon as being a stream of negatively charged particles. He named the negatively charged particle an electron.

Sir Ernest Rutherford noted the presence of positive charges in the nucleus of the atom.

Later Rutherford and Neils Bohr gave us the first picture of the atom as a miniature solar system with electrons on planetary rings surrounding the nucleus which contained the protons.

Mendeleeff produced the first accepted chart of the atoms arranged according to weight. It contained certain irregularities. These irregularities were necessary to enable Mendeleeff to place related elements in the same family.

Moseley, through complex experimental work and calculations, showed that the arrangements of the atoms should be on a basis of atomic number.

The atomic number is the number of electrons found in the planetary rings of the element. These are matched by an equal number of protons in the nucleus.

It was noted that there was a discrepancy between the atomic weight and the number of protons in all cases except hydrogen.

Chadwick discovered the neutron, a neutral particle of great density which was found in the nucleus of the atom. It was found that the total number of protons plus the total number of neutrons gave the atomic weight of the element.

QUESTIONS

Part A

1: What was the strange phenomenon that Crookes noted during his famous experiment?
2: To whom did Crookes appeal for an explanation?
3: What was the explanation concerning this "ray of light"?
4: What is the nature of an electron?
5: What was the basis of Rutherford's reasoning that led to his discovery of the proton?
6: What caused scientists to think there were other components present in the atom in addition to protons and neutrons?
7: What is the nature of a neutron?
8: Who discovered the neutron?
9: Of what is the total weight of an atom composed?
10: Who gave us our first arrangement of the atoms according to weight?
11: Examine the chart of the atoms and note the three pairs of elements that are out of order in respect to weight.
12: What is the value of the atomic number when you are asked to draw a diagram of an atom?
13: How do we calculate the number of neutrons present in the nucleus?
14: How many electrons are found in a complete K ring?
15: How many electrons are found in a complete L ring?
16: Draw a diagram representing each of the following atoms: hydrogen, helium, barium, carbon, nitrogen, calcium.
17: In what two ways does our diagram of the atom fail to give a true picture?
18: What was Robert Millikan's contribution to the study of the atom?

Part B

19: Which halogen (fluorine, chlorine, bromine, or iodine) will show the greatest attraction for the electron? Why?

20: Lithium, sodium, and potassium all lose one electron when forming a compound with nonmetals. Which will lose its electron most readily? Why?

21: The chemical properties of elements are a periodic function of their atomic numbers. Explain.

22: Draw diagrams of the atoms lithium, sodium, and potassium, and from the diagrams indicate why they might have similar properties.

23: Chlorine, argon, and potassium appear on the chart as elements 17, 18, 19. One is an active nonmetal, one is an inert element, and one is an active metal. Explain how this can be.

IN THE LABORATORY

To study differences between mixtures and compounds

Place one spoonful of iron powder on a sheet of paper. Examine it as to color, appearance (metallic or nonmetallic), and hardness. Place a few grains in a test tube half-filled with carbon disulphide and see if it will dissolve. With the aid of a magnifying lens examine it and note to the best of your ability if each granule is the same as the next granule. Test with a magnet. Note results.

On a second piece of paper place a spoonful of powdered sulphur. Examine it carefully in the same manner that you examined the iron. Note the results.

Take half of the iron and half of the sulphur and place them on a third piece of paper.

Using a splint or toothpick thoroughly mix the iron and the sulphur. Examine the mixture and note the results.

Do the properties of the original elements disappear?

Take the remaining half of the iron and place it in a crucible. Take the remaining half of the sulphur and mix it thoroughly with the iron. Using a clay triangle to support the crucible, place it on a ring stand and heat it slowly at first and then strongly, using a Bunsen burner as a source of heat. Continue to heat it until the contents of the crucible glow brightly. Allow it to cool. Examine the contents of the crucible following the same
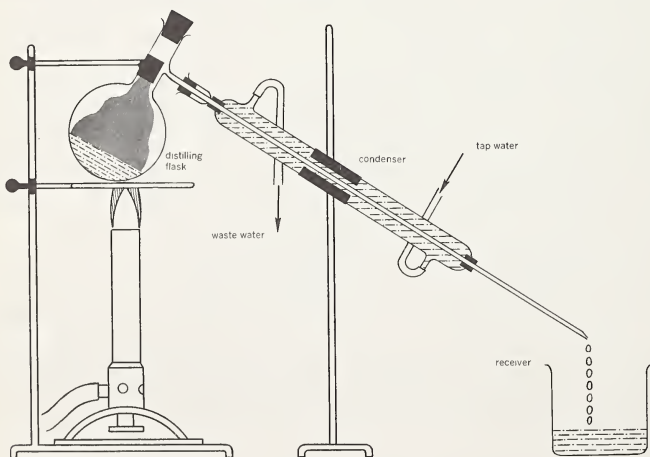

Fig. 4-9. Distillation of water using a Liebig condenser.

scheme that you used in examining the original elements. Make a chart like the one below and note the results.

Do the contents of the crucible resemble the original elements in any major property?

Why should we expect this result?

What conclusion may be drawn concerning the properties of a compound as compared with the properties of the elements that compose the compound?

DISTILLATION OF WATER

Natural water taken from rivers and wells, while it may be safe for drinking, is not pure water. Explain why we would not expect to find pure water in nature.

For use in the laboratory the chemist requires pure water. The process of purifying natural water is known as distillation and consists of boiling the impure water and collecting the condensed steam. The process itself is quite simple, but if much distilled water is required it is more convenient to order it through bottling works, drug stores, or chemical supply houses.

To obtain pure water from this process one must use absolutely clean apparatus, and the impurities must have a higher boiling point than that of water.

In the laboratory the impure water is usually boiled in a flask and the steam is directed through a Liebig condenser. The condenser consists of a central tube through which the steam passes. This tube is surrounded by a second tube through which cold water is in constant circulation (see Fig. 4-9).

DISTILLING WATER IN THE LABORATORY

Completely dissolve a spoonful of salt in tap water. Pour the solution into a flask and arrange the apparatus as shown in Fig. 4-10. Using a Bunsen burner as a source of heat, bring the solution of salt and water to a boil. Allow the boiling to continue until 5 or 10 cc. of water have been collected in the test tube which is known as a receiving vessel. The liquid (pure water in this case) in the

Property	Iron	Sulphur	Mixture	Compound
Color				
Appearance				
Hardness				
Solubility				
Similarity of granules				
Effect of magnet				
Any other properties noted				

receiving vessel is known as the distillate. Taste the distillate. Is it free of salt?*

In some cases, where the impurity has a boiling point close to the boiling point of the distillate, the process has to be repeated several times.
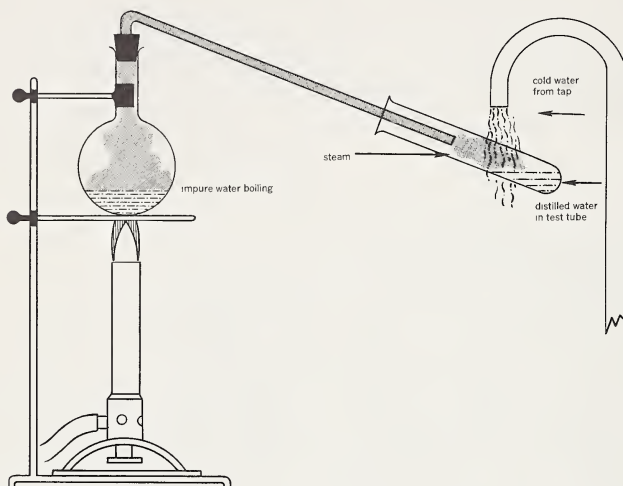
Fig. 4-10. A simple apparatus for distilling water. Such a setup may be used for student experimentation.

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* Taste is not considered a scientific test and the student should not taste any chemicals or reagents unless directions indicate that it is safe to do so. Later in the course, certain scientific tests for simple substances will be introduced.

SECTION 2

An Introduction to Chemistry

UNIT 3 - THE LANGUAGE OF CHEMISTRY

Unfortunately many students pick up a chemistry text, and because they have some interest in the subject they quickly look at the pictures, glance at the introduction, pass over the fundamentals, and start reading about the middle of the book. Naturally they run into trouble. They come face to face with chemical symbols and equations and quickly decide that chemistry is not for them.

Actually, if the study of any shorthand is made from the start, it should not prove difficult. A word at a time, an idea at a time, and the whole plan should become plain.

We use shorthand in our daily conversation to reduce the number of words and to save time. A friend may ask you what a certain show is all about, and you may take half an hour replying in a detailed narrative. Or if you are in a hurry you may simply say, "Boy meets girl, lives happy ever after."

In the world of chemistry certain basic information can be expressed in short forms. In fact, whole industrial processes can be outlined in a few lines of symbols and equations. To try and understand these with a few minutes study would be impossible, and to grasp them it is necessary to start at the beginning.

An introduction to symbols, valence, compounds, and equations follows in the next two chapters.

CHAPTER **5** SYMBOLS AND FORMULAE: THE SHORTHAND OF CHEMISTRY

We have already seen a sample of the chemist's shorthand in Chapter 3 when we considered the subject of chemical symbols. Symbols represent atoms of elements which are the smallest particles of matter that take part in a chemical reaction. In addition to merely representing the name of the element, we learned that the symbol also stands for a definite weight of the element represented. This is known as the atomic weight. (Review Chapter 3, page 34.)

The compound is represented by placing two or more symbols or radicals together. For example, sodium chloride, the principal constituent of table salt, is a compound composed of 1 atom of sodium and 1 atom of chlorine. The *symbol for the atom* of sodium is Na and the *symbol for the atom* of chlorine is Cl. The *formula for the molecule* of sodium chloride is NaCl.

In the early years when compounds were first being studied, the chemist was at a loss to explain why sodium and chlorine, or for that matter any other two elements, should unite to form a new substance. The best they could offer was not an explanation but merely a statement that sodium and chlorine were attracted to each other. Later they represented this attraction by showing the symbol as possessing a hook with which it linked itself to the hook of another element:

Na ⌐⊃⊂⌐Cl   Na⌐π⌐Cl
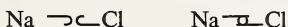
This became shortened to the form Na— —Cl and the compound Na—Cl. The attraction was still unexplained.

By the middle of the nineteenth century the Swedish chemist, Svante Arrhenius, offered one of the most remarkable theories that had been suggested in the history of science. A modification of it is known today as the *Arrhenius' theory of ionization.* While this course does not concern itself with his theory, one of the implications of it is important to us.

Arrhenius suggested that while the atoms themselves were neutral, when a compound was formed there was some form of an electrical attraction between the component parts of the compound. He demonstrated that certain compounds could dissociate in water solution. Sodium chloride was one of these compounds and, instead of producing atoms of sodium and atoms of chlorine both of which are neutral electrically, the dissociation produced positively charged particles of sodium, represented as Na^+ and negatively charged particles of chlorine represented as Cl^-. These charged particles Arrhenius named *ions.*

In the early twentieth century, with the discovery of the electron, proton, and neutron (review page 41) and the formation of the electron theory, it was realized that Arrhenius had been years ahead of his time in his thinking and that neutral atoms did assume charges when they formed compounds and retained those charges when the compound dissociated. It was also noted that *the atoms became charged by gaining or by losing electrons in a chemical reaction. The protons and neutrons remain unchanged.*

Accordingly, sodium as an element is neutral electrically and is represented as $Na°$. When it becomes charged it does so by losing an electron as represented:

$$Na° - e \rightarrow Na^+$$

55
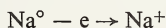
In a like manner chlorine, neutral as an element, gains an electron to become the ion of chlorine with a negative charge as represented:

$$Cl^\circ + e \rightarrow Cl^-$$
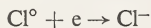
WHY DO ELEMENTS LOSE AND GAIN ELECTRONS?

Earlier in our study we read of elements being active, inactive, or inert. By the careful observation of many experiments and the interpretation of the data based on the atomic numbers of the elements, scientists found that varying degrees of activity depended upon, among other things, the atomic structure of the element with special reference to the number of electrons on the outside ring.

The above may be illustrated by a look at the Periodic Chart of the atoms that appears on the inside back cover of this text. The inert elements (Group O) you will note have complete outer rings of electrons. We say they are completely *stable*. Helium with 2 electrons on its outer ring, neon with 8, argon with 8, and so on, do not form compounds of any kind.*

Scientists also noted that other elements with less than complete outer rings showed varying degrees of activity, and would give

away or accept electrons from other elements or would share electrons with other elements in order to complete the outer rings of their structure and thus achieve stability. The complete outer ring is known as a *stable octet.*

Let us see how this applies to the compound sodium chloride. From the atomic weights and the atomic numbers of sodium and chlorine we can construct diagrams representing these two elements (Fig. 5-1). You will notice that sodium has 1 electron on its outer ring and that chlorine has 7. Therefore neither element is as stable as it would be in a compound. Sodium could do one of two things. It could give away 1 electron or it could accept 7. Chlorine could accept 1 or give away 7. It is obvious that the simpler way of gaining a stable octet would be for sodium to give away the 1 electron and chlorine to accept it to form the compound sodium chloride as illustrated in Fig. 5-2.

The compound itself is neutral and the two elements that compose it are now both stable.

When a compound is formed by the transfer of electrons such as is illustrated in Fig. 5-2, we say that the compound is an *electrovalent compound* because it consists of electrically charged parts.

Many of our compounds are not electrovalent. Instead of giving or receiving electrons the elements that form the compounds share pairs of electrons. Each pair is made up of one electron from one element and one electron from another element. Compounds formed in this way are said to be covalent, and the attraction that holds the elements together as a compound is known as a covalent bond.

Earlier it was stated that certain of the element gases formed molecules containing 2 atoms. In such molecules we also find examples of covalent bonds. For example, consider the atom and the molecule of hydrogen. From the atomic number and weight we can represent the atom of hydrogen as in Fig. 5-3.

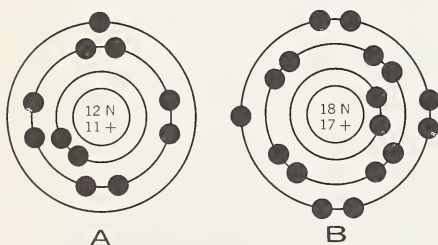The atom appears, however, with an incomplete ring, and so is unstable. As the

Fig. 5-1. *A.* The element sodium. *B.* The element chlorine.

12 N
11 +

A

18 N
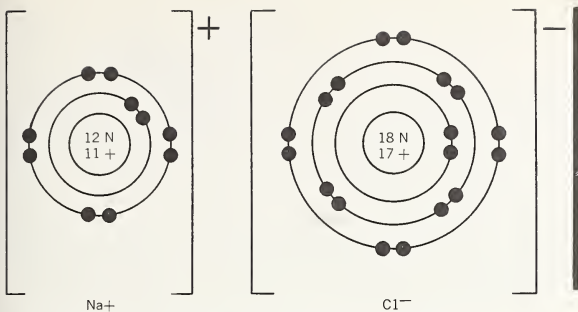17 +

B

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* See footnote on p. 31.

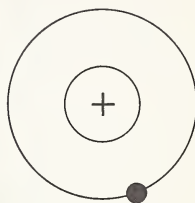Fig. 5-2. The electrovalent compound of sodium chloride, NaCl.

Fig. 5-3. The hydrogen atom.

complete ring in this case consists of two electrons, the only way the atom can achieve stability in the element form is to unite with a second atom of hydrogen which will give the diatomic (2 atoms) molecule as illustrated in Fig. 5-4.

It is thought that the electrons revolve around both protons (Fig. 5-5).*

ELECTROVALENT COMPOUNDS

We have studied the changes that take place when sodium, which loses 1 electron, unites with chlorine, which accepts 1 electron, to form an electrovalent compound. Now let us study the changes that take place when magnesium unites with sulphur to form the compound, magnesium sulphide. By studying the weight and number of these two

elements we note that magnesium has 2 electrons to give away and sulphur will accept 2. The reaction may be represented as shown in Fig. 5-6.

Because the reaction takes place with 1 atom of magnesium uniting with 1 atom of sulphur the formula, therefore, will be MgS or $Mg^{++}S^{--}$.

Building formula would be quite a simple matter if all the combinations of the elements would fit as easily as we have seen in the above illustrations. But let us see what happens when we exchange the chlorine and the sulphur in the two compounds and have the sodium of the sodium chloride unite with the sulphur and the magnesium of the magnesium sulphide unite with the chlorine.

Because sodium has 1 electron to lose and the sulphur has to accept 2 electrons to gain
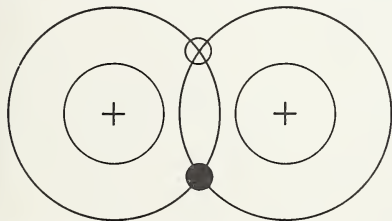
Fig. 5-4. The molecule of hydrogen. The black dot represents one electron belonging to one of the atoms while the open dot represents the electron belonging to the other atom. The nucleus of each atom is a single proton.
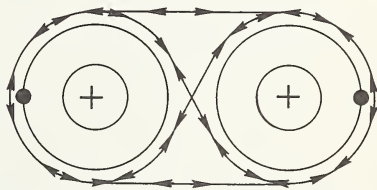

Fig. 5-5. The covalent molecule of hydrogen.

stability, you will readily understand that it will require 2 atoms of sodium and 1 atom of sulphur to form the compound known as sodium sulphide (Fig. 5-7).

* See footnote on page 44.
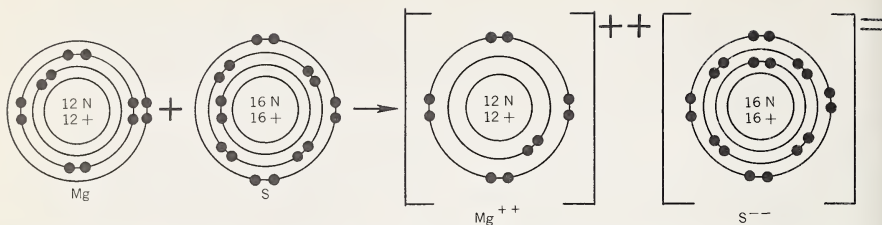
Fig. 5-6. During the reaction of magnesium with sulphur the magnesium loses two electrons which are accepted by the sulphur. It is an electrovalent compound.

When the union is complete we could represent the compound of sodium sulphide in the manner illustrated by Fig. 5-8.

On the other hand, when magnesium and chlorine unite, the magnesium has 2 electrons to give away but the atom of chlorine can accept only 1 electron. The final compound formed by this reaction known as magnesium chloride must use a second atom of chlorine which accepts the second electron which magnesium must lose. Magnesium chloride may be represented as seen in Fig. 5-9.

To indicate, in the final formula of each of the two compounds mentioned above, that there are 2 atoms of one element we place a small subscript after the symbol of the element doubled. Thus, sodium sulphide

with 2 atoms of sodium becomes Na_2S and magnesium chloride with 2 atoms of chlorine becomes $MgCl_2$.

The above compounds are electrovalent compounds.

COVALENT COMPOUNDS

As we read above, there are certain compounds the atoms of which share rather than accept or donate electrons. The stability attained by sharing electrons giving a stable octet (or duet in the case of hydrogen) can be seen by examining the diagrammatic representation of a molecule of methane gas which has the formula CH_4.

Carbon has four electrons on its outer ring while hydrogen has only one.

Fig. 5-7. Sodium reacts with sulphur.

Fig. 5-8. The compound sodium sulphide. Na_2S is electrovalent and is often written as $2Na^+ + S^{--}$

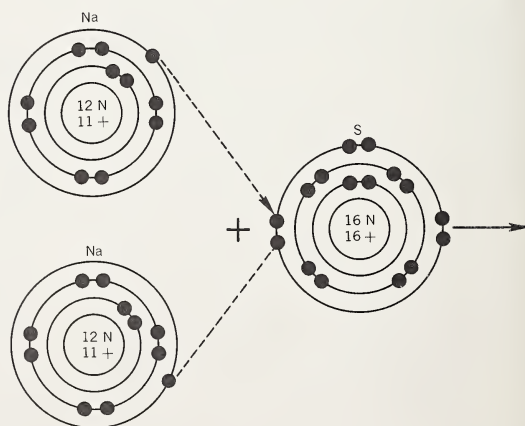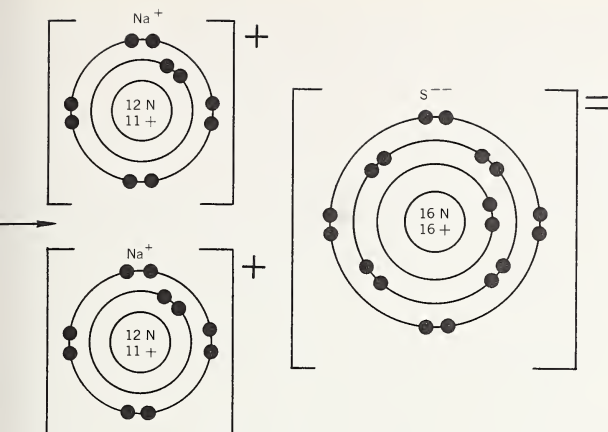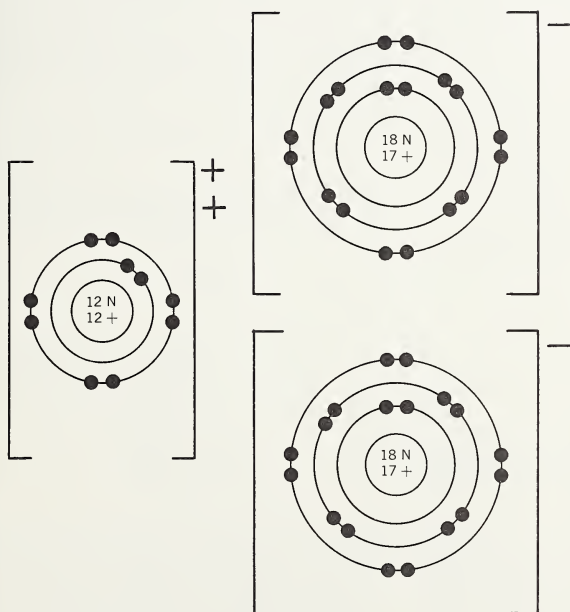**Fig. 5-9.** The electrovalent compound magnesium chloride. $MgCl_2$ or $Mg^{++}2Cl^-$

In Fig. 5-10 the electrons belonging to the hydrogen atoms have been represented by circles (○) while those of the carbon have been represented by dots (●). There is no difference between the two electrons and they are represented in this fashion merely to help you interpret the diagram.

Another method of representing such a compound can be seen in Fig. 5-11.

TRANSFER OF ELECTRONS AS WELL AS SHARING ELECTRONS MAY BE REPRESENTED BY "BONDS"

It would indeed be a complicated process if every time a chemist wished to show the structure of a compound he had to draw the electron arrangement. To avoid such a situation he uses "bonds" to represent the loss or the gain of single electrons or the sharing of a pair of electrons.

Thus because sodium loses a single electron it can be represented by the symbol for sodium with a bond, as Na—, and because chlorine accepts a single electron it can be represented as the symbol for chlorine with a bond, as —Cl. The formula for the compound, therefore, is Na^+Cl^-.*

In the case of magnesium losing 2 single electrons, we have 2 bonds showing $Mg\big\langle$,

and magnesium chloride would be $Mg\big\langle{Cl \atop Cl}$

or ($Mg^{++}2Cl^-$).

The structural formula for methane is usually shown as

$$H-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{C}}-H$$
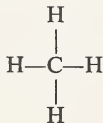
In the case of methane it is not shown with the carbon and the hydrogen carrying plus
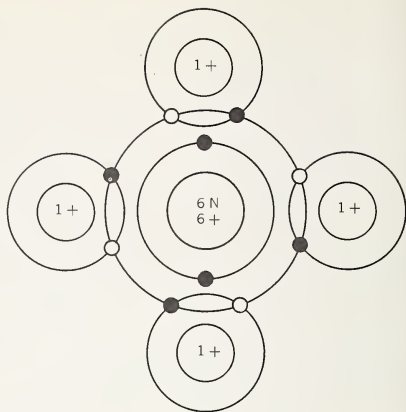
Fig. 5-10. The covalent molecule of methane CH_4.

and minus charges, as the compound is covalent and shares pairs of electrons rather than electrovalent which involves the transfer of electrons. Therefore, in the case of methane, you must remember that each bond represents a pair of shared electrons. The pair is made up of one electron from the hydrogen and one electron from the carbon repeated four times.

VALENCE

You will note that in the above examples *each element has a definite number of electrons*
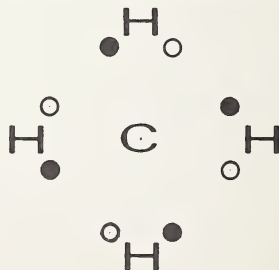
Fig. 5-11. A diagrammatic representation of the covalent compound, methane. The black dots represent the electrons on the outer ring of the carbon while the open circles represent the electrons supplied by the four hydrogen atoms.

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* As you advance in modern chemistry you will find that the use of the bond is reserved for the covalent molecule such as in methane (CH_4). The electrovalent compound is represented by the symbols showing the charge as Na^+Cl^-, $Mg^{++}2Cl^-$, etc. As soon as you master the idea of valence it would be well to follow this system.

to give away, to accept, or to share when it forms a compound. This number of electrons which it gives away or accepts or shares is known as the valence of the element.

METALS AND NONMETALS

In an earlier chapter we discussed the physical properties of two groups of elements known as the metals and the nonmetals. Now we are in a position to look at the important chemical property that distinguishes one group from the other.

Those elements that give up electrons to achieve a stable octet are called metals while those elements that accept electrons are nonmetals.

Because metals lose electrons when they form compounds they develop a positive charge. Nonmetals accept electrons and therefore are negative when they form compounds. Because of these facts, even though the element is in an uncombined state, we speak of it as having a valence of $+1$ or $+2$ or -1 or -2, as the case may be. Actually, in the free state it is not exhibiting any valence and should be spoken of as having a valence of zero while it is in the free state. When we assign a valence to it in the free state, as we usually do, we really mean that it has a potential valence of $+1$, or -1, and so on.

As would naturally be expected, those elements that have less than one-half of the possible electrons on the outside ring have a greater tendency to lose electrons than to accept them, and these, therefore, form the metal group. They appear to the left of the "staircase line" on the chart on the back endpaper in this text. The nonmetals appear on the right-hand side.

The accompanying table lists some of the more common elements together with the symbol of each and its valence expressed as a plus ($+$) or a minus ($-$) depending on its behavior as a metal or a nonmetal. These elements, divided as metals and nonmetals, are also referred to as cations (ionized metals) and as anions (ionized nonmetals).

MULTIPLE VALENCE

For various reasons some elements have more than one valence. Some elements even display more than one valence in a single compound.

An example of multiple valence results from internal changes in the electronic structure of the element. According to your Periodic Classification of the Atoms the ring structure of the atom of copper is 2, 8, 18, 1. This indicates that copper has a valence of $+1$ because it will lose one electron to give it a stable structure. Under certain conditions, however, the ring structure of copper becomes 2, 8, 17, 2, and the element changes from monovalent to divalent.

Other common elements that display this type of multiple valence are mercury, with valences of $+1$ and $+2$; iron, with valences

	Monovalent			Divalent		Trivalent	
Metals	Lithium	Li^+		Barium	Ba^{++}	Aluminum .	Al^{+++}
	Sodium	Na^+		Calcium	Ca^{++}	Iron (ic) . . .	Fe^{+++}
	Potassium	K^+		Copper (ic) . .	Cu^{++}		
	Copper (ous)	Cu^+		Iron (ous) . . .	Fe^{++}		
	Silver	Ag^+		Magnesium . .	Mg^{++}		
	Mercury (ous)	Hg^+		Mercury (ic) . .	Hg^{++}		
				Zinc	Zn^{++}		
Nonmetals	Bromine	Br^-		Oxygen	O^{--}	Nitrogen . .	N^{---}
	Chlorine	Cl^-		Sulphur	S^{--}	Phosphorus .	P^{---}
	Fluorine	F^-					
	Iodine	I^-					

of $+2$ and $+3$; and tin, with valences of $+2$ and $+4$.

These differences are, of course, shown in the formula of any compound involving the element. Copper with a valence of $+1$ is united with oxygen with a valence of -2 in the ratio of 2 atoms of copper to 1 of oxygen, and the formula is therefore Cu_2O.

The formula would appear as:

$$Cu \diagdown$$
$$\qquad O$$
$$Cu \diagup$$

When copper is displaying a valence of $+2$ the formula of the oxide would be CuO and its structural formula would appear as:

$$Cu{=}O$$

Both of the above compounds are oxides of copper.* But they are different compounds with different properties. It is, therefore, necessary that we distinguish between them when we name them. This is achieved by adding a suffix to the name of the metal in each case. To indicate the lower valence of the metal the suffix *ous* is added, and to indicate the higher valence of the metal the suffix *ic* is added to the metal. The compound Cu_2O is therefore known as cuprous oxide, and the compound CuO is called cupric oxide.

The same rule applies to the other compounds which are made up of elements with multiple valence. Iron (ferrum) with a valence of 2 forms ferrous compounds and with a valence of 3 forms ferric compounds. Mercury with a valence of 1 forms mercurous compounds and with a valence of 2 forms mercuric compounds.
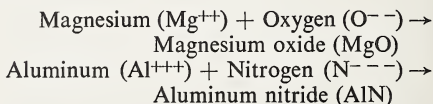
WRITING OF FORMULAE

Before a formula can be written, the valence

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* Many modern texts are beginning to introduce a new system of nomenclature to indicate the valences of metals in compounds. Cuprous oxide, for example, because copper is exhibiting a valence of one, is indicated as copper I oxide; and cupric oxide, where copper has a valence of two, is shown as copper II oxide. In a like manner, the iron oxides are shown as iron II oxide and iron III oxide.

of each element in the compound must be known.

We shall look at some simple compounds made up of two elements. Such compounds are known as **binary** compounds. The fact that a compound is a binary compound is indicated in the name by adding the suffix -*ide* after the name of the nonmetal in the compound. Thus, because sodium and chlorine unite to form a binary compound, the compound is known as sodium chloride. A compound of copper and chlorine, if copper has a valence of $+1$, would be known as cuprous chloride. The -*ous* indicates the low valence of copper. If copper had a valence of $+2$ the compound would be called cupric chloride, the suffix -*ic* indicating the high valence of copper.

When we were studying the structure of the atom and how compounds were formed by the transfer of electrons, we noted that there was little difficulty in forming compounds by the union of metals and nonmetals which have the same valence. For example, sodium chloride is composed of sodium with a valence of $+1$ and chlorine with a valence of -1 to give us Na^+Cl^- or $NaCl$. Other examples are:

Magnesium (Mg^{++}) + Oxygen $(O^{--}) \rightarrow$
Magnesium oxide (MgO)
Aluminum (Al^{+++}) + Nitrogen $(N^{---}) \rightarrow$
Aluminum nitride (AlN)

More frequently, however, we will wish to indicate formulae where the valence of each component part is not the same.

There are short methods which enable the student to write the formula quickly and accurately, but before we consider these let us see how a formula is constructed, using bonds to represent gain and loss of electrons.

As an example, let us take the problem of writing the formula of aluminum oxide.

Aluminum is a metal, and from its atomic structure we see that it has three electrons on its outside ring. To gain stability it will lose these three electrons. Oxygen is a nonmetal. It has six electrons on its outside ring and will accept two electrons to give it a stable octet.

Aluminum, therefore, may be represented as Al, and oxygen as O. Placing the elements side by side with bonds matching, aluminum and oxygen would appear as

$$Al\diagdown\hspace{-0.5em}>O.$$
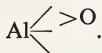
As indicated by the bonds, such a combination would leave aluminum with one electron on its outside ring and therefore unstable. To overcome this, another atom of oxygen is added and the structure appears as

$$Al\diagdown\hspace{-0.5em}\begin{array}{c}>O\\>O.\end{array}$$

The diagram now shows oxygen as being unstable, and to supply more electrons a second atom of aluminum is added.

$$\begin{array}{c}Al\diagdown\hspace{-0.5em}>O\\>O\\Al\diagdown\end{array}$$

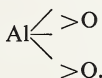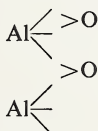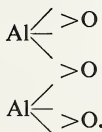While oxygen became stable aluminum now has two electrons to lose which, of course, can be accepted by a third atom of oxygen as illustrated

$$\begin{array}{c}Al\diagdown\hspace{-0.5em}>O\\>O\\Al\diagdown\hspace{-0.5em}>O.\end{array}$$

It is obvious that the above indicates the formula of aluminum oxide as Al_2O_3.

This method of constructing formulae may be applied to any of the common compounds that you will meet, and if it is studied carefully, remembering that each bond represents a single electron gained or lost or a pair of electrons shared, you should quickly master the subject of valence.

Once an understanding of the structure of the compound has been mastered we will realize that in reality we are uniting the two elements by using the lowest common denominator of their respective valences.

The same result can be reached by using the valence of the metal as the subscript of the nonmetal of the compound, and using the valence of the nonmetal as the subscript of the metal as seen with reference to aluminum oxide:

$$Al_2 \quad O_3$$

RADICALS

Certain groups of atoms form combinations that are quite stable and act as if they were a single element in chemical reactions. These groups are known as radicals. They possess a definite valence.

In your study of chemistry you will use many of these groups, and you will in advanced courses learn of their structure and how they may be broken into their component parts. In this course, however, you will use only a few radicals and their names and valences should be memorized.

When you write formulae containing radicals it is necessary to use parentheses around the radical if the subscript is greater than 1. This indicates that the whole radical is taken two or three times, as the case may be, rather than the last element of it only. For example, because zinc has a valence of $+2$ and the sulphate radical has a valence of -2, the formula of zinc sulphate is $ZnSO_4$ but in the case of zinc nitrate, the nitrate radical has a valence of 1 and will therefore have to be taken twice. Hence the formula of zinc nitrate will be $Zn(NO_3)_2$.

While valences should be memorized and through use they should come to the mind without thought, there may be times when the valence of a particular element is momentarily forgotten. In such a situation you should try to remember a formula in which the element in question is found, and from it you will be able to determine the forgotten valence. For example, if you are asked to write the formula of sodium sulphate and you know the valence of sodium is $+1$ but you have forgotten the valence of the sulphate radical, you should try to remember the formula of some other compound containing that radical. Possibly the formula of sulphuric acid as H_2SO_4

comes to mind. You know the valence of hydrogen is $+1$ and as two hydrogen atoms are in the formula, the valence of the sulphate radical must be -2. Therefore, the formula of the compound in the original question, sodium sulphate, will be Na_2SO_4.

TABLE OF COMMON RADICALS

Radicals may be divided into two groups depending on the charge that they carry. Those that are negatively charged are known as **anions**. Sometimes anion radicals are known as acid radicals because they can unite with hydrogen to form acids. In some texts they are called nonmetallic radicals. The only positively charged radical discussed in this course is the ammonium radical NH_4^+. It is known as a **cation**.

	Monovalent		Divalent		Trivalent	
Cation	Ammonium	NH_4^+				
Anions	Bicarbonate	$(HCO_3)^-$	Carbonate	$(CO_3)^{--}$	Phosphate	$(PO_4)^{---}$
	Bisulphate	$(HSO_4)^-$	Sulphate	$(SO_4)^{--}$	Phosphite	$(PO_3)^{---}$
	Bisulphite	$(HSO_3)^-$				
	Chlorate	$(ClO_3)^-$				
	Hydroxide	$(OH)^-$				
	Nitrate	$(NO_3)^-$				
	Nitrite	$(NO_2)^-$				

THINGS TO REMEMBER

The symbol of the element represents its atom. It also represents 1 atomic weight of the atom.

The formula is the representation of two or more atoms in chemical union. Examples of formulae are as follows:
1. The formula of the diatomic molecule, as O_2, H_2, etc.
2. The formula of the radical, as SO_4^{--}, NO_3^-, etc.
3. The formula of the compound, as $NaCl$, KOH, $Al_2(SO_4)_3$.

Compounds are formed by the chemical union of atoms that have lost or accepted electrons or that have shared electrons.

Metals are those elements that form compounds by losing electrons. This results in a stable octet on the outer ring of the element. They appear on the left side of the heavy "staircase line" on the chart of the classification of elements on the back endpaper of this text.

Nonmetals are those elements that form compounds by gaining electrons. This results in a stable octet on the outer ring of the element. They appear on the right side of the staircase line referred to above.

The number of electrons which an element will gain, lose, or share in forming a compound is known as the valence of the element in that compound.

Radicals are groups of atoms that act as a unit in a chemical reaction. Each radical has its own valence.

QUESTIONS

Part A

1: What does the symbol represent beside the name of an element?
2: What does the formula represent?
3: Who was the first chemist to suggest the theory that electrical charges played a part in the forming of compounds and in their dissociation?
4: What is the theory known as today?
5: Today we say that sodium chloride dissociates to the sodium ion with a positive charge and to the chlorine ion with a negative charge. Why do we consider sodium as charged positively and chlorine as charged negatively?
6: How do we account for the charges on each of these ions?
7: Why do we say calcium loses 2 electrons when forming a compound?
8: Why do we say oxygen gains 2 electrons when forming a compound?
9: What do you understand by the term "stable octet" in relation to atomic structure?
10: Explain the difference between "neutrality" and "stability" of an element.
11: How does a nonmetal achieve stability in forming a compound?
12: How does a metal achieve stability in forming a compound?
13: Compounds may be formed by electron transfer or the giving or receiving of electrons. What other method do you know by which compounds may be formed?
14: Draw a diagram representing a chlorine molecule.
15: Suggest a reason why metals lose electrons to achieve stability while nonmetals gain electrons to achieve the same result.
16: What relationship has the ability to gain or lose electrons to the valence of the element? Illustrate your answer by reference to (a) hydrogen and (b) chlorine.
17: Why is helium said to have a valence of zero?
18: Why is it necessary to know the valence of an element before writing the formula of a compound of the element?
19: Account for the fact that copper has a valence of $+1$ and may also have a valence of $+2$ depending upon certain conditions.
20: What do the *-ous* and *-ic* suffixes refer to when used in the naming of iron compounds?
21: What suffix does the chemist use to indicate a binary compound?
22: Write the formula and name the compound in each of the following instances: a compound made of (a) calcium and chlorine, (b) sodium and bromine, (c) aluminum and nitrogen, (d) potassium and sulphur.
23: What is a radical?
24: Write the formula and indicate the valence using $+$ or $-$ signs for each of the following radicals: (a) phosphate, (b) sulphate, (c) carbonate, (d) bicarbonate, (e) bisulphate, (f) chlorate, (g) ammonium.

Part B

25: Draw the structural formula of aluminum sulphate. Consider the radical as a single unit SO_4.
26: Name and give the structural and chemical formulae of compounds made up of the metals and radicals named below. In each case consider the radical as a single unit and do not try to show its structure.
(a) sodium and the sulphate radical
(b) copper and the carbonate radical
(c) zinc and the nitrate radical
(d) potassium and the phosphate radical
(e) iron (valence $+3$) and the sulphate radical
(f) potassium and the chlorate radical

27: If you wished to write a certain formula that required you to know the valence of zinc and you found that you had forgotten this, suggest a method that you might use to reason it out from other formulae that you happened to remember.

28: The formula of table sugar is $C_{12}H_{22}O_{11}$. What does this formula tell you about the composition of sugar both as to elements in it and the relative quantities of those elements? What is its gram molecular weight?

29: Explain what the "bond" stands for in the structural formula of (a) an electrovalent compound and (b) a covalent compound.

30: Explain the meaning of the following symbols and formulae:
$$N, \quad N_2, \quad 2N_2.$$

31: Make a list of all the properties that you can think of that will show the difference between the metal copper and the nonmetal sulphur.

IN THE LABORATORY

In Chapter 5 we studied the theory of electron transfer between metals and non-metals. Magnesium, for example, will unite with oxygen to form a new compound, magnesium oxide. Let us observe some of these reactions in the laboratory.

To observe the differences between elements and their compounds

Copper. Examine a short length (2 inches) of copper wire that has been removed from its insulating cover. Note its properties as you can observe them, its color, its luster, its hardness. Now, holding the wire with forceps, heat it in the flame of a Bunsen burner until you are sure that no further change is taking place. Allow it to cool and note the appearance of the wire comparing it with a piece that has not been heated. Enter your findings in the chart below.

Magnesium. Again note the properties of the element, examining it as to appearance, color, pliability, etc. When you heat it be careful that you do not burn yourself, as it will ignite and burn with an extremely hot flame giving off a brilliant light. Examine the results. Make a chart in your notebooks and enter your findings.

Lead. Repeat the experiment using a strip of lead that has been filed so that a bright surface is exposed.

To observe if there is any change in weight when a metal is heated in air

Place a small amount of clean steel wool in a porcelain crucible and balance it against a known weight on a set of chemical balances. (NOTE: it is possible that the steel wool may be oily. If such is the case rinse it in a beaker of carbon tetrachloride and dry it. Be careful not to breathe the fumes of the carbon tetrachloride as they are poisonous.)

Metal heated	Properties before heating	Properties after heating
Copper		
Magnesium		
Lead		
Aluminum		
Other metals tried		

Using a clay triangle to support the crucible, place it on a ring stand and heat it slowly at first and then strongly, using a Bunsen flame as a heat source.

After heating the crucible and contents for about ten minutes allow it to cool and replace it on the balance.

If the balance is delicate enough you will be able to note the difference in weight. Account for the difference remembering that the iron was heated in the presence of air (oxygen).

The final goal of the chemist is the end-point of the chemical reaction. It will give him the product for which he has searched or the breakdown of the compound that he has analysed. His method of attaining his desired result is the story of the chemical changes that have taken place. The chemist uses what is called a chemical equation to represent this process in a concise form.

You will notice that the equation is the description of a reaction that has already taken place and not a prophecy of what the chemist expects or hopes will happen. In order that this idea does not escape us, in our elementary study of equations we will first conduct an experiment and then to the best of our ability decide what has taken place before we write the equation of the reaction.

EXPERIMENT

6-1. A simple reaction → a simple equation

Take a small amount of mercuric oxide* (HgO), the amount you can pile on a ten-cent piece, and place it in a hard-glass (pyrex) test tube. Using a clamp, place the tube in the flame of a Bunsen burner and heat it quickly.
Almost immediately you will note that the color of the oxide has changed. A few moments later you will notice that a silvery-looking substance has condensed near the top of the tube. Further, if you insert a glowing splint into the top of the test tube, the splint will burst into flame. In our work we consider this last reaction of the splint bursting into flame as a test for the presence of a high concentration of oxygen. If you continue heating the tube

you will see that eventually the complete sample of the mercuric oxide has disintegrated. After the tube has cooled, use a splint to remove the silvery substance from the test tube and collect it on a watch glass. Examine it carefully. From its appearance and weight and the fact that it came from mercuric oxide, we would say it was mercury. While this is not a scientific test, the answer is correct and we will accept it.
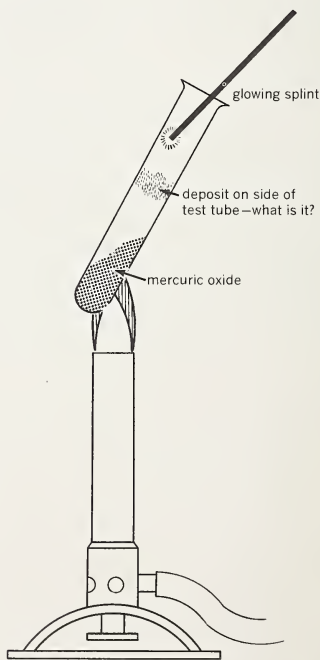

Fig. 6-1. Heating the red oxide of mercury (HgO) in a pyrex test tube.

*See footnote, p. 90

Using the shorthand of the chemist let us now show the action of heat on mercuric oxide.

mercuric oxide + heat delivers mercury and oxygen

Substituting the formula for the names of the various substances used and produced and substituting an arrow (→) for the word "delivers," we have the following:

$$HgO \rightarrow Hg + O$$
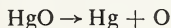
This is known as the simple or the skeleton equation. You will notice that heat or any other form of energy used in the reaction is omitted from the equation.

Earlier in this course we learned that the molecule was the smallest particle of any substance that could be identified as that substance, and therefore we must work with molecules and not atoms. We also learned that certain gases at ordinary temperatures and pressures contained 2 atoms to the molecule and that oxygen was one of these. Therefore the representation of oxygen by "O" is incorrect in the equation which should read:

$$HgO \rightarrow Hg + O_2 \uparrow *$$
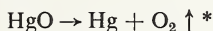
You will notice that in a compound we may have an atom of oxygen, but in the free state oxygen must be represented as a molecule.

From your knowledge of an equation in mathematics or from the very definition of the word itself you know that both sides of the equation must be of equal value.

The last representation is therefore not a true equation because on the left side we have 1 atom of oxygen while on the right side we have 1 molecule of oxygen composed of 2 atoms. This must be corrected.

Without giving the matter any thought you might suggest that we change the formula of

mercuric oxide from HgO to HgO_2. This would appear to satisfy our needs. But such a suggestion overlooks the fact that the composition of matter is constant. That is, mercuric oxide is HgO and always will be HgO according to the law of definite proportions.

The only way left, then, to get a second atom of oxygen is to take a second molecule of mercuric oxide. The reaction may then be represented as:

$$2HgO \rightarrow Hg + O_2 \uparrow$$

But this is still not an equation as the balance of the mercury is wrong. On the left side we have 2 atoms of mercury and only 1 on the right side.

This error is simply one of mathematics and may be corrected by recording a second unit of mercury on the right side to give us the balanced equation

$$2HgO \rightarrow 2Hg + O_2 \uparrow$$

You will notice while we cannot change the subscripts which are used to give us the formula of a compound, we are at liberty to change the numbers in front of the formula for they simply indicate the amounts of the reacting materials. These numbers are known as coefficients.

When the coefficient 2 is placed in front of a formula it indicates that the whole formula is doubled. The coefficient is written the full size of the capital letter. The subscript is written in a smaller size and is placed below and after the element or radical to which it refers.

For example:

$2HgO$ represents 2 molecules of mercuric oxide each composed of 1 atom of mercury and 1 atom of oxygen. $2Hg$ represents 2 atoms of mercury and it also may represent 2 molecules of mercury as mercury is a monatomic molecule. O represents 1 atom of oxygen while $2O$ represents 2 atoms. O_2 represents 1 molecule of oxygen, as indicated above, while $2O_2$ represents 2 molecules. Once we have used coefficients and subscripts for a short while, difficulties should disappear. One fact that must be

---

* When writing an equation in which one of the reacting substances or one of the reacting products is a gas, an arrow pointed upward is usually placed immediately after the formula: e.g., $H_2 \uparrow$, $CO_2 \uparrow$, $O_2 \uparrow$. In a like manner if the substance is a precipitate, the formula of the precipitate is usually followed by an arrow pointing downward: e.g., $AgCl \downarrow$, $BaSO_4 \downarrow$.

remembered, however, is that *subscripts of a chemical formula are constant and must never be changed for correction of quantities.*

THE WRITING OF EQUATIONS

Before any chemical reaction may be written in the form of an equation, there are certain things that we must know. We must know that a chemical reaction is possible and we must know the composition and names of the reacting substances and products as established by chemical tests. Many students believe that the writing of chemical equations is simply writing chemical formulae and then seeing what possible combinations can be made from them. Nothing could be further from the truth. The following will illustrate this fact.

EXPERIMENT

6-2. *To write the equation for the reaction of dilute hydrochloric acid and magnesium*

Place about one inch of magnesium ribbon in a test tube and add about 15 cc. of dilute hydrochloric acid. Note that the bubbling of the reacting materials indicates that a gas is being evolved. Test the gas coming from the
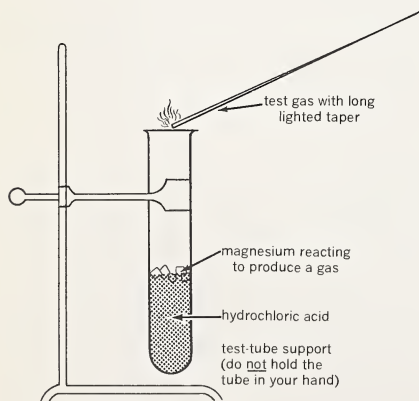

test gas with long lighted taper

magnesium reacting to produce a gas

hydrochloric acid

test-tube support (do <u>not</u> hold the tube in your hand)

Fig. 6-2. Action of magnesium and hydrochloric acid.

reaction with a burning splint. (Caution: be careful that the test tube is directed away from yourself and not directed toward any other person as the gas is explosive.) There will be a sharp "bark" as the gas in the tube explodes. Under these conditions and using the substances that we did, this indicates to us that hydrogen gas has been released. Further tests, which we are not yet able to perform, would indicate that magnesium chloride ($MgCl_2$) has been formed and remains in solution in the test tube.

If the above reaction is represented in the form of an equation it would appear:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \uparrow$$
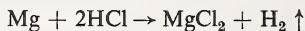
Now let us consider another problem that appears to be quite similar in nature. Let us discuss the possible reaction between copper and hydrochloric acid.

On first thought you would likely say that the reaction would be quite similar to the one described above, because all that we have done is substituted copper for magnesium. You might even go as far as to suggest that the equation could be written without further authority, such as might be gained by reading or actually trying the reaction in the laboratory.

As part of your science training, however, try this reaction by taking a small piece of copper, or a penny, place it in a test tube, and add acid similar to that used in the above experiment with magnesium. Note your result and discuss it with your teacher.

From this experiment you should see the necessity of proving that a reaction is possible and of knowing what is formed thereby, before attempting to write the equation that describes it.

Once a chemist knows that a reaction is possible he proceeds to write the equation following a definite plan. The procedure is outlined as follows:

1. The names of all reacting substances and the names of all products must be known.

2. The formula of each reacting substance and each product must be known. These

must be arranged in the form of a skeleton equation.

3. Examine the right side of the equation and make sure that everything that has been supplied by the left side has been used.

4. Any changes on the right side may have affected the left side. Examine it and make sure that there is sufficient material to supply the right side.

5. The processes described in steps 3 and 4 may have to be repeated until both sides are in balance.

6. Check the balance.

If the above steps are followed, the balancing of an equation becomes a mathematical procedure, and guesswork on the part of the student should be eliminated. There is no need for trial methods being used.

EXPERIMENT

6-3. To check the results of heating potassium chlorate with manganese dioxide and to describe the reaction in the form of an equation

Place about 10 grams of potassium chlorate on a sheet of clean paper. Add to this a quantity of manganese dioxide equal to about *one-third the bulk of the chlorate. Thoroughly mix the two compounds. Slide the mixture into a test tube, clamping the tube in a position that is almost horizontal. Tap the tube to distribute the mixture evenly over two-thirds of its length. Use a single-hole rubber stopper and arrange a delivery tube to a water trough for collection of the gas as seen in Fig. 6-3.*

Heat the contents of the tube by a Bunsen burner allowing the air to bubble through the water at first. Later collect a tube of the gas and test it with a glowing splint. Because of the fact that the splint burst into flame, we assume that there was a concentration of free oxygen in the tube that must have come as a result of the reaction in the heated tube.

At this stage of your studies in chemistry you are not able to test the products that remain in the original test tube. From tests that have been conducted, however, we have found that two substances remain. The first is potassium chloride (KCl), apparently the potassium chlorate ($KClO_3$) has lost its oxygen, and the second is the original manganese dioxide (MnO_2), apparently unchanged. Further tests on the part of chemists have shown that while this substance did not actually enter the reaction, it did speed the reaction. Such a substance, which changes the speed of a reaction, is
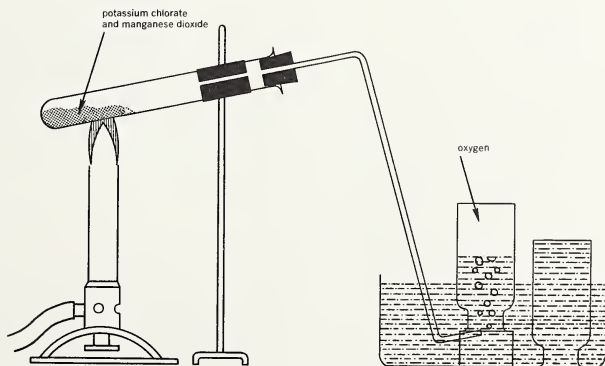

potassium chlorate
and manganese dioxide

oxygen

Fig. 6-3. The laboratory preparation of oxygen.

known as a **catalyst**.* It does not need to be shown in the written equation.

Let us consider the representation of this reaction in the form of a chemical equation. We shall follow the points as outlined on page 70 step by step.

1. The names of reacting substances and products.

   Potassium chlorate →
           potassium chloride + oxygen.

2. The formula of each substance.

   $KClO_3 \rightarrow KCl + O_2$ (Remember the oxygen must appear as a molecule.)

3. Examine the right side of the equation to determine if all material supplied by the left side has been used.
   a. The symbols K and Cl appear on both sides of the equation in equal amounts.
   b. On the right side we have 1 molecule of oxygen consisting of 2 atoms, but on the left side we have 3 atoms of oxygen. It is apparent that we have not completely used the supply of oxygen. This could be balanced if we were permitted to multiply the oxygen on the right side by $1\frac{1}{2}$, but as fractions are not acceptable as coefficients we must double it, giving a coefficient of 3. The equation then appears:

   $$KClO_3 \rightarrow KCl + 3O_2$$

4. Examine the left side to make sure that we have sufficient material. With the raising of the oxygen to 3 molecules (composed of 6 atoms) on the right side, we shall have to adjust the oxygen on the left side, which originally consists of 3 atoms. This will require it to be doubled. As we cannot change the subscript, it will have to be done through the coefficient, and the equation becomes:

   $$2KClO_3 \rightarrow KCl + 3O_2$$

5. Again examine the right side.
   The doubling of the $KClO_3$ means we now have 2 possible molecules of KCl after the oxygen has been given up, but only one shows on the right side. This must be doubled. The equation now appears:

   $$2KClO_3 \rightarrow 2KCl + 3O_2$$
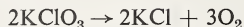
6. Check the balance.
   In checking the balance remember that the coefficient number refers to the whole formula that follows, but the subscript refers only to the atom that precedes it or to the radical within parentheses that precedes it. Total values are obtained by multiplying the coefficient by the subscript. Remember, if no subscript shows, it is taken as 1.

In the case of the KCl, where it remained as a group it would have been possible to consider it as a single unit in the check as:

$$KCl \quad 2 \times 1 = 2 \quad 2 \times 1 = 2$$

Thus we see that the equation is in balance.

You will notice in the above that the equation shows the basic reaction but omits certain detail. It does not show the presence of any required catalyst. It does not show whether the substances are dry or whether they are in solution. It does not show the concentration of the reacting materials, or the conditions under which they react. While all this information is of importance to the chemist, most of it comes with experience. In the meantime, the student will have to watch for instructions in carrying out the suggested demonstrations.

TYPES OF CHEMICAL REACTIONS

Most of the reactions that we shall be studying in this course can be classified as one of four types, or possibly a combination of two of the types which follow.

Type 1. Direct Combination Sometimes Called Direct Synthesis–

---

* Do not be misled by the above statement concerning the role of manganese dioxide as a catalyst. Many students who meet MnO_2 for the first time in the preparation of oxygen consider that it is always a catalyst. This is incorrect. Although it is a catalyst in this reaction, in other reactions it is an active compound and in no way a catalyst.

Element	Quantity on left side Coefficient | Subscript | Total	Quantity on right side Coefficient | Subscript | Total
K	2 × 1 = 2	2 × 1 = 2
Cl	2 × 1 = 2	2 × 1 = 2
O	2 × 3 = 6	3 × 2 = 6

EXPERIMENT

*(Caution: As hydrogen is highly explosive in air it is suggested that this reaction be carried out using a Kipp generator. It should be demonstrated by the teacher.)**

6-4. To demonstrate the direct synthesis of water by burning hydrogen in air

Using a hydrogen generator charged with zinc and hydrochloric acid dilute (HCl), produce sufficient hydrogen gas to fill a large test tube. The gas may be collected by the downward displacement of air.

Place a cover glass over the open end of the inverted test tube and move it a safe distance from the generator (at least 10 feet) before turning it with the open end up and igniting the gas with a burning taper.

If the hydrogen is impure, that is, mixed with air, there will be a sharp "bark" as it explodes. If it is pure it will simply burn with a non-luminous flame. In both cases you will notice that water has formed on the inside of the test tube.

The above represents the direct synthesis of water by the burning of hydrogen in air. The active part of the air is oxygen. The reaction may be represented by the chemical equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$
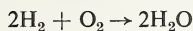
ASSIGNMENT

Starting with a simple word equation describing this synthesis show how we arrive

* Because of the explosive nature of hydrogen, you are warned not to light the gas at the generator until the properties and the correct method of handling hydrogen have been studied in a further course. One test tube filled with the gas and ignited as described above should be sufficient to demonstrate the reaction.

at the balanced equation following the steps as suggested on page 70 of the text.

Type 2. Simple Decomposition (this is the reverse of direct synthesis) – You have, no doubt, in your elementary school science passed an electric current through water that has been acidified with a small amount of sulphuric acid. The result that you observed was the water breaking down into its component parts, hydrogen and oxygen. This is an example of simple decomposition. The equation describing the reaction is:

$$2H_2O \rightarrow 2H_2 + O_2$$
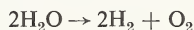
ASSIGNMENT

Starting with a simple word equation show how we arrive at the final balanced equation following the steps as indicated on page 70.
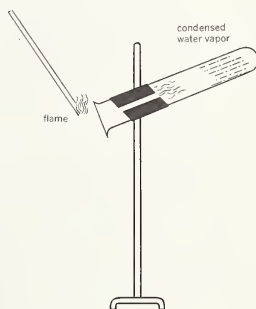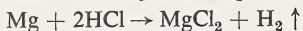

Fig. 6-4. Water is the result of hydrogen burning in air.

Type 3. Simple Replacement or Simple Substitution – We have already had an illustration of this type of reaction on page 70. When magnesium reacts with hydrochloric

acid, the magnesium replaces the hydrogen of the acid according to the equation:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2 \uparrow$$

While we found that at least one other metal does not react in the same way as magnesium, there are a great number of metals that do replace hydrogen from acids.

ASSIGNMENT

Take small samples of zinc, iron, and silver with hydrochloric acid and treat each in the same way that you treated the magnesium on page 70. If there is any indication of a reaction taking place, test the gas given off with a burning splint. Write the equations for the reactions.

Type 4. Double Replacement or Double Decomposition – You will have observed that most of the compounds that we have studied so far are made up of the union of metals and nonmetals or by the union of metallic and nonmetallic radicals.

In the fourth type of reaction the metals of two compounds simply exchange. To illustrate this let A and B represent two metals and x and y represent two nonmetals. If the original compounds were Ax and By the final result of the reaction would be Ay and Bx.

In some reactions the evidence that a chemical change has taken place may be seen by the evolution of a gas or precipitate.

EXPERIMENT

To illustrate a double decomposition reaction

Dissolve 1 gram of barium chloride in half a test tube of water. Dissolve 1 gram of sodium sulphate in half a test tube of water. Pour the contents of test tube 1 into test tube 2 and note the change. The white precipitate that you observe is barium sulphate (BaSO₄). Remaining in solution is sodium chloride.

ASSIGNMENT

Write the balanced equation for the reaction. Many of the reactions are combinations of two or more of the above types.

EXPERIMENT

**To illustrate a reaction that is a combination of two of the above types**

Drop a small piece of limestone (calcium carbonate, CaCO₃) into a test tube of sulphuric acid. Observe the evolution of a gas and test it with a burning splint.

The fact that the burning splint was extinguished indicates that the gas was not oxygen which, while it does not burn, does support combustion. It also indicates, as the gas did not burn, that it was not hydrogen. (While hydrogen does not support combustion, it does burn.)

On the authority of those who have tested and identified this gas, you must accept it as carbon dioxide.

The question immediately arises: Where did it come from?

The answer does not appear immediately, for if we examine the reacting substances we see that the expected reaction would be one of a double replacement type. It could be expressed in the equation

$$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2CO_3$$
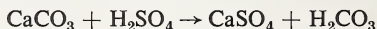
Actually the reaction is a double replacement but there is also a secondary reaction that we must consider. The carbonic acid formed (H_2CO_3) is a very unstable compound and, unless kept under pressure, loses its form as such a compound by changing to water and carbon dioxide gas:

$$H_2CO_3 \rightarrow H_2O + CO_2 \uparrow$$

The complete reaction, therefore, is a combination of a double replacement reaction and a simple decomposition reaction and may be represented by one equation which includes both:

$$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2 \uparrow$$

ANOTHER USE OF THE EQUATION

In addition to providing the chemist with a valuable method of describing a reaction, the equation may be used as a basis for establishing the quantitative relationships of the reacting substances.

When he is discussing the substances that react and the products formed by the reaction, the chemist is interested not only in the "what" but also in the "how much."

He may use the equation to give him the answer to the quantities involved because of two facts.

1. The atomic weights mentioned earlier (review Chapter 3, page 34) give us a relationship between the weights of the various atoms.
2. The law of definite proportions (review Chapter 3, page 33) tells us that compounds are always composed in definite proportions by weight.

An application of this may be seen in the reaction, with which we are already familiar, of iron being heated with sulphur.

Let us apply the atomic weights to the elements and the compound as represented by the equation:

$$Fe + S \rightarrow FeS$$
$$56 \quad 32 \quad 88$$
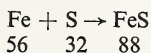
If we conducted the experiment very carefully, weighing the materials with great accuracy, after being certain that the reaction was complete we would find that 56 grams of iron would unite with 32 grams of sulphur to form 88 grams of iron sulphide.

But suppose that we had 112 grams of iron and all the sulphur we required to complete the reaction, what weight of iron sulphide would be produced?

If you examine the relationship between 112 and 56 you will see that the larger weight is double the smaller. Therefore, we would expect to get twice the amount of iron sulphide. This calculation may be written as a mathematical expression in the form of a ratio, $112 : 56 = x : 88$. This is, of course, taken directly from the balanced equation:

$$112 \qquad x$$
$$Fe + S \rightarrow FeS$$
$$56 \qquad 88$$
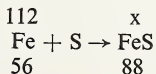
x being equal to 176 grams of iron sulphide.

It is not necessary to have the components of our ratio on opposite sides of the chemical equation, as the relationship established

through the atomic weights applies to all of the reacting substances as well as the products.

For example, in the same reaction the question could have read: What weight of sulphur would be required to react completely with 112 grams of iron?

Applying the atomic weights (here in the case of the elements they are also the molecular weights) to the balanced equation

$$112 \quad x$$
$$Fe + S \rightarrow FeS$$
$$56 \quad 32$$
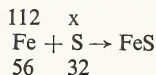
we find x = 64 grams of sulphur.

You note that we place the molecular weights, as taken from the periodic table, below the symbols of the elements to which they apply, and we place the given weight and the symbol of the unknown (x) over the elements to which they apply. Once the ratio is established we solve as a simple algebraic ratio.

We usually speak of the weight of an element as its molecular weight. The molecular weight of iron in the above equation is 56 as indicated in the table on the back endpaper of this text. As has been mentioned, this is a relative weight and may be described as being a gram molecular weight or an ounce molecular weight, depending on the unit of the given weight in the problem. The unit of weight, of course, must be consistent throughout the problem and the solution.

If we were using oxygen or any element which has a diatomic molecule, we would have to multiply the atomic weight by 2. Thus oxygen with an atomic weight of 16 having the molecule O_2 would have a molecular weight of $16 \times 2 = 32$.

If the above equation had called for 2 molecules of iron and been written as 2Fe, the weight placed under the 2Fe would have been $56 \times 2 = 112$.

On the right side of the equation we have the compound, iron sulphide, expressed as a formula made up of 2 atoms. To find the weight of this substance we add the weights of the atoms that make up the compound. The atom of iron has the same weight as the weight of a molecule of that element which

is 56. The same applies to the atom of sulphur with a weight of 32. The total weight of the compound is therefore 88. This may be expressed as grams, ounces, and so on.

If the result of the reaction had been 2FeS we would, of course, have had to double the weight placed under it in the equation.

In this course it is acceptable to refer to the weight of iron sulphide as the molecular weight of 88, or if it is to be expressed in grams, as the gram molecular weight of 88 grams.

As you advance in chemistry, however, you will find that the weight of an electrovalent compound is referred to as the formula weight rather than the molecular weight. The weight of a covalent compound is always referred to as the molecular weight. While the reason for this difference in nomenclature is not discussed in this course it would be well to become accustomed to its use whenever you recognize a compound as being electrovalent or covalent.

Example 1: *To find the percentage composition of a compound from its formula*

If you had 10 pounds of fruit made up of 6 pounds of apples and 4 pounds of oranges, you could calculate the percentage weight of the oranges by a simple computation. Your reasoning would be that with the total weight being 10 pounds, 4 pounds of which are oranges, the percentage of oranges would be

$$\frac{4}{10} \times 100 = 40 \text{ per cent oranges by weight.}$$

Because the assigned weights of the atoms are relative, we can carry out similar calculations in computing the percentage by weight of any element or group of elements in a compound.

Example 2: *To calculate the percentage of ferric oxide* (Fe_2O_3)

1 Note atomic weights of iron and oxygen:

$$Fe = 56 \qquad O = 16$$

Formula weight of the compound:

2 atoms of Fe =	$2 \times 56 =$	112
3 atoms of oxygen	3×16	48

$$\text{Total formula weight} \quad 160$$

2. Percentage of iron present:

$$\frac{112}{160} \times 100 = 70 \text{ per cent}$$

3. Percentage of oxygen present:

$$\frac{48}{160} \times 100 = 30 \text{ per cent}$$

$$100 \text{ per cent}$$

The same procedure may be applied to a group of elements in a compound.

Example 3: *To find the percentage of water in a crystal of barium chloride which has the formula* $BaCl_2 \cdot 2H_2O$*

Note the atomic weights present:

1 atom of barium	137
2 atoms of chlorine 35.5×2	71
2 molecules of water 2×18	36
Total formula weight	244

The percentage of water, therefore, would be:

$$\frac{36}{244} \times 100 = 14.7 \text{ per cent}$$

---

* In calculating weights in chemistry the dot stands for an addition sign.

THINGS TO REMEMBER

Equations are records of a chemical reaction written in an abbreviated form.

The left side of an equation shows the materials entering the reaction.

The right side of the equation shows the products of the reaction.

Molecular formulae must be used at all times.

The equation must be balanced.

Subscripts used in any chemical formula of a given compound are constant and must never be changed for that compound.

Coefficients are the numerals placed in front of the formula and represent the number of molecules of the substance used in the reaction.

Equations are balanced by examining the right side of the equation making sure that all the material supplied by the left side has been used. The left side is then examined to make sure that there is sufficient material for the needs of the right side.

There are four main types of reactions:
1. Direct synthesis.
2. Simple decomposition.
3. Simple replacement.
4. Double decomposition or replacement.

Equations are the basis of computing quantitative relationships. The ratio is established by placing the formula weight below the corresponding formula in the equation. The given weight is placed above the corresponding formula. The unknown is placed over the corresponding formula. The ratio is solved by algebraic methods.
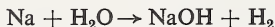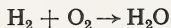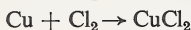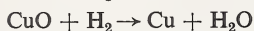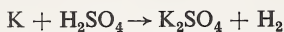
QUESTIONS

Part A

1: In Chapter 6 you were introduced to the study of equations. In each case, before the equation was presented the reaction was always demonstrated. What important fact about equations was being emphasized by this procedure?
2: Write the balanced equation showing the result of heating mercuric oxide.
3: Why is it necessary to balance the equation?
4: Why would it be incorrect to avoid the difficulty by writing mercuric oxide as having the formula HgO_2?
5: Expand your answer to question 4 to show that we must never change subscripts in a formula.
6: Outline from memory the steps followed in writing and balancing an equation. Illustrate your answer by using the following reaction:

Sodium hydroxide + sulphuric acid → sodium sulphate + water.

7: What is a catalyst?
8: Is it possible for a substance to be a catalyst in one reaction and an active substance in another?
9: Name and give an example of four types of chemical reactions mentioned in this chapter.
10: Explain the statement: "Equations may be used for establishing quantitative relationships in chemical reactions." Illustrate your answer with reference to the reaction of

potassium chlorate heated in the presence of the catalyst manganese dioxide to deliver potassium chloride and oxygen.

Part B

11: Balance, if necessary, the following equations and tell which type of reaction each illustrates.

$$K + H_2SO_4 \rightarrow K_2SO_4 + H_2$$
$$CaCO_3 \rightarrow CaO + CO_2$$
$$CuO + H_2 \rightarrow Cu + H_2O$$
$$Cu + Cl_2 \rightarrow CuCl_2$$
$$H_2 + O_2 \rightarrow H_2O$$
$$Na + H_2O \rightarrow NaOH + H_2$$

12: How does the composition of a molecule of oxygen differ from the molecule of mercuric oxide with respect to oxygen?

13: Why is a chemical equation called an equation?

14: What facts are not shown by the chemical equation?

15: What must we know before we are able to write an equation?

16: By experiment we may show that 14 grams of iron unite with 8 grams of sulphur. How many grams of iron sulphide will be produced? Write the equation and arrange the given and formula weights to show 14 grams of iron and 8 grams of sulphur reacting.

17: What weight of sulphur will be needed to unite with 11.2 grams of iron so that there will be no iron left over at the end of the reaction? What will be the weight of the iron sulphide formed?

18: What weight of potassium chlorate must be used to produce 8 grams of oxygen?

19: What weight of mercuric oxide would be required to produce 8 grams of oxygen?

20: If we passed dry hydrogen over heated copper oxide the hydrogen would unite with the oxygen to form water and leave metallic copper as illustrated by the equation:

$$CuO + H_2 \rightarrow Cu + H_2O$$
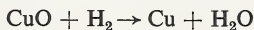
If we started with 80 grams of copper oxide what weight of hydrogen would be required to result in a complete reaction? What weight of water would be formed? What weight of copper would be produced?

21: What is the percentage composition of sulphuric acid?

22: Crystallized barium chloride is represented by the formula $BaCl_2 \cdot 2H_2O$. What is the percentage of water present in the molecule?

23: How many grams of $CaCO_3$ reacting with sulphuric acid would be required to deliver 110 grams of CO_2?

24: What weight of carbon dioxide would be produced if an excess of calcium carbonate reacted with 392 grams of a solution of sulphuric acid that contained 75 per cent H_2SO_4?

25: What weight of zinc would be required to displace all the hydrogen from 182.5 grams of hydrochloric acid? If we used iron instead of zinc to replace the hydrogen what weight of iron would be required? (When you are writing your equation the salt produced will be $FeCl_2$.) If we used aluminum instead of zinc what weight of aluminum would be required?

IN THE LABORATORY

To determine experimentally the percentage of water present in crystals of barium chloride

Weigh a clean dry crucible together with its cover to the nearest centigram. Note the weight.

Take sufficient crystallized barium chloride that will about half fill the crucible, place it in the crucible, and weigh the crucible contents and cover. Note the weight.

Place the covered crucible and contents on a ring stand. It may be mounted by means of a pipe stem triangle. Heat the crucible slowly at first and then quite strongly for about 10 minutes. Remove the cover and continue to heat for another 5 minutes. Allow the crucible to cool.

Reweigh the crucible, contents, and cover. Note the weight.

Reheat the open crucible for another 5 minutes and reweigh.

Repeat this process until the weight of the crucible, cover, and contents is constant. Note the final constant weight.

Fill in the chart below and calculate the percentage of water that was present in the original crystals.

Weight of crucible, cover, and crystals =
Weight of crucible and cover (empty) =
Weight of crystals of barium chloride =
Constant weight of dry barium chloride =
Water of hydration (driven off by heat) =

Calculate the percentage of water of hydration in barium chloride crystals.

NOTE: If you do not have barium chloride crystals in your laboratory you may do a similar experiment using crystalized copper sulphate ($CuSO_4 \cdot 5H_2O$).

Calculate the percentage of error in your experiment by comparing the answer you found experimentally to the theoretical answer you calculated in question 22 in this chapter.

If time permits, ask your teacher for a sample of a crystal unknown to you. After determining the percentage of water of hydration check your results with your teacher.

It would appear to the student that the chemist is most happy when he is sorting, arranging, and classifying everything from atoms to compounds.

In Unit Two we have read some of the scientists' theories and noted a few fundamental facts concerning the states, composition, and the structure of matter.

In this unit we will study some of the substances in which the chemist is interested. We will see how he uses the knowledge gained from his study of pure science to give man new substances as well as new uses for old substances. This study is known as applied science.

Unit Four discusses the element oxygen, one of the most important of all the elements. In spite of its importance nothing was known about it until after 1770. Today it is known as the gas of a thousand uses.

A study of three classifications of compounds: acids, bases, and salts completes the unit. Each classification is made up of many compounds and each group is valuable to man.

Do not make the mistake of thinking that these three classifications exhaust the groups of substances found in the field of chemistry. As you continue your study you will find many new classes and subclasses of substances in both inorganic and organic chemistry.

CHAPTER 7 OXYGEN: THE WORLD'S MOST ABUNDANT ELEMENT

For centuries after man had begun to wonder about the composition of matter, the great sea of air in which he lived still remained one of the deep mysteries. Certain things were known about it, some materials would burn in it, it was necessary in order to support life of plants and animals. But apart from these simple facts even the scientist of the Middle Ages knew very little about the actual composition of the atmosphere.

Today practically every student of elementary science is well aware of the fact that the atmosphere in which we live is made up of two principal gases, oxygen and nitrogen, plus small amounts (about 1 per cent of the total) of a number of other gases.

It must be remembered that the gases that compose the atmosphere are mixed together and not chemically combined.

We cannot name any single element as being the most important. There are, however, some elements of vital importance to us in our life on this earth, and oxygen can be classed as one of these.

Because of the great abundance of oxygen in the atmosphere (one-fifth of the volume of the air is made up of this element) it comes in contact with most of the substances that we find in our earth. Because of the great activity of the element oxygen, it unites, slowly or rapidly as the case may be, with many of these substances. The product of such a chemical union is called an oxide.

When we take into account the oxygen of the air as well as the oxygen found in natural compounds, we consider that 49.8 per cent of the earth's crust including the atmosphere is made up of oxygen either in the free state or compounded with one or more elements.

HISTORY OF OXYGEN

Joseph Priestley, an English chemist and clergyman, is generally given credit for the discovery of oxygen. According to his diary of August 1, 1774, he heated mercuric oxide (known to him as red precipitate). The method he used was to invert in a dish of mercury a glass vessel filled with mercury and a little mercuric oxide. To supply the energy necessary for the reaction, he used the heat of the sun's rays, concentrated by a lens (Fig. 7-1).

As the red precipitate disappeared he noted that a gas was being formed, and after
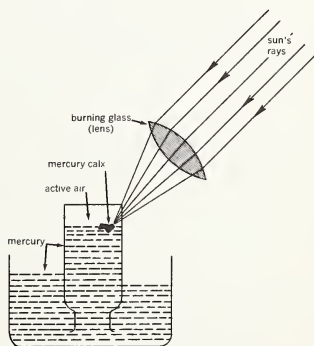

Fig. 7-1. Priestley's preparation of oxygen. In Priestley's day mercuric oxide was known as mercury Calx and oxygen was called "active air."

collecting a sample he found that a candle would burn in the new gas much more brightly than it would burn in air.

Previous to this time K. W. Scheele, a Swedish apothecary who conducted chemical research in his attic bedroom, had produced the same gas in 1771. As Scheele did not publish the account of his discovery until 1777, credit for the discovery is given to Priestley.

By 1777 the brilliant French chemist, Antoine Lavoisier, became interested in the new gas that had been produced. He decided to study its properties and establish its relationship to air. By means of a charcoal fire he heated a volume of mercury in a closed retort. The neck of the retort passed through another quantity of mercury and opened into a bell-jar containing air (see Fig. 7-2).

The mercury over the fire was heated for 12 days. During this period Lavoisier noticed that a red powder was forming on the surface of the mercury that was being heated. He also noticed that as the powder formed on the mercury, the quantity of air in the bell-jar diminished as indicated by the rising level of the mercury in the bell-jar (Fig. 7-2).

No more red powder formed and the volume of the air did not diminish after the volume of air in the jar had been reduced by one-fifth.

On the reheating of the red powder there was given off an active gas with properties similar to those of the gas described by Priestley. Its volume matched the volume of the loss of air in the bell-jar.

To the active gas, in which a candle would burn more vigorously than in air, Lavoisier gave the name "oxygen" which means "acid producer." Later, when the true nature of acids became known, it was realized that the term "oxygen" as applied to the gas was incorrect. It has still remained, however, to this day.

THE PREPARATION OF OXYGEN

The study of the preparation of oxygen must be divided into two sections, (1) the preparation of the gas in the laboratory and (2) the preparation of the gas for use in industry. In each case we must, of course, start with a
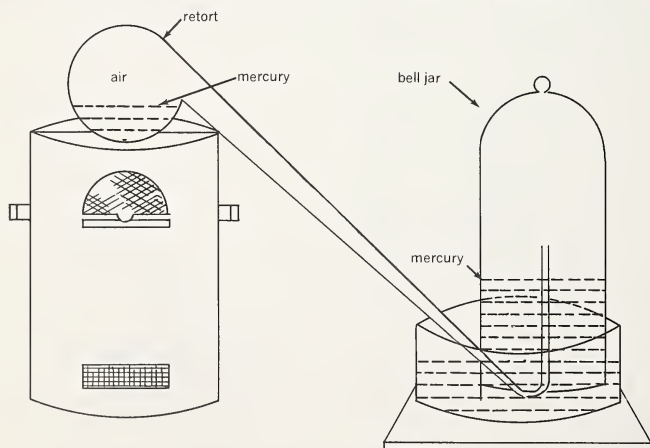

Fig. 7-2. Experiment carried out by Lavoisier which proved that one-fifth of the air had united with the mercury to form mercuric oxide.

substance that will yield oxygen. In the case of industry, the substance must be relatively inexpensive and the process as simple as possible to keep costs low. In the case of the laboratory, where small amounts are used, ease of production is the important factor even if it involves the use of a relatively expensive substance as a source of the element. Ease of production also implies the use of simple laboratory equipment.

The Preparation of Oxygen in the Laboratory – While we stress convenience and ease of handling materials as important factors to take into consideration when we prepare materials in the laboratory, nevertheless, if two materials are both handled easily, we naturally choose the one that produces the product more cheaply.

There are three common methods of preparing oxygen in the laboratory.

1. *By the heating of mercuric oxide.* This is the method by which Priestley discovered the gas. We are familiar with the reaction from demonstrations carried out earlier in this course (see page 68).
2. *By the heating of potassium chlorate using manganese dioxide as a catalyst.* At this point it would be advisable to review the demonstration described in Chapter 6, page 71. As this is possibly the most popular method of producing oxygen for the laboratory since we can obtain a good volume of the gas comparatively cheaply, the process should be mastered thoroughly and the equation well known.
3. *By the action of water on sodium peroxide.* If water is added a few drops at a time to sodium peroxide, oxygen is delivered in accordance with the following equation:

$$2Na_2O_2 + H_2O \rightarrow 4NaOH + O_2 \uparrow$$
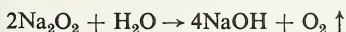
Caution: One must be careful that water is added in small amounts (a few drops at a time), otherwise so much oxygen will be produced in such a short space of time that the apparatus will be damaged and the experimenter may be injured.

One must also be careful that any sodium peroxide left after the reaction is placed in the sink and water added slowly until the reaction is complete. Do not throw it in waste baskets, or any place where it will come in contact with inflammable material.

There are other methods by which oxygen may be prepared in the laboratory, and if you are interested you could find them in a more advanced chemistry textbook.

The Preparation of Oxygen in Industry – Many millions of cubic feet of oxygen are produced in Canada each year for uses in industry and for medical purposes.

As pointed out earlier, one of the most important factors to ensure a successful industrial preparation of any substance is an abundant source of raw material that may be obtained at low cost. No source could be cheaper and more abundant for the preparation of oxygen than the air itself.

As oxygen forms roughly 20 per cent of the air by volume and as it is in an uncombined state because air is not a compound but a mixture, the whole problem of obtaining oxygen from the atmosphere is simply a method of separating the gas from the other gases in air by a mechanical means.

The industrial production of oxygen takes advantage of several significant facts concerning the behavior of gases. First, gases cool rapidly when they expand in volume. You may demonstrate this for yourself by noting that the automobile tire and the air inside the tire may be so warm after traveling over hot highways on a summer day that your hand would be burned if you tried to keep it on the tire casing for even a few seconds. Yet, if you pressed the valve in such a way as to release some of the air from the tire, the rush of expanding air feels cold as it strikes your hand. Second, air can be liquefied by cooling it to a temperature of about −200°C. And third, the oxygen and nitrogen that form the mixture known as air have different boiling points and can be separated by means of allowing one gas to boil off while the other is retained as a liquid.

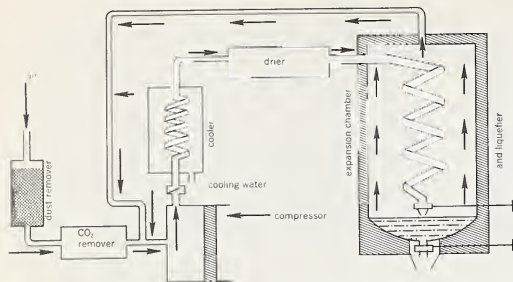Fig. 7-3 shows the basic steps in the production of industrial oxygen.

Fig. 7-3. The industrial preparation of oxygen and nitrogen.

Air is taken from the atmosphere and the dust is removed. The small percentage of carbon dioxide present in all air is then removed by a chemical process. The air is then compressed until the pressure is about 200 atmospheres. (By 200 atmospheres we mean 200 times the pressure that the air exerts in our atmosphere or $200 \times 14.7 = 2940$ lb. per sq. inch.) This results in the temperature of the air being raised due to the compression process. (You have noticed that a bicycle pump becomes quite warm when it is used to compress the air for a tire.) The warm compressed air is piped through a cooling system and then through a "drier" where the moisture is removed. After further cooling, the air is then forced through a small opening and allowed to expand in a chamber. This drops the temperature further. Any gas cooled sufficiently to be a liquid remains in the chamber and is led back to the compressor and the process is repeated.

The liquid nitrogen is then allowed to boil off. When clear of nitrogen the liquid oxygen is allowed to boil and the gaseous oxygen is compressed in tanks for delivery.

There is another method which supplies a small amount of oxygen for industry and which may also be used to supply oxygen for the laboratory. In this case, the oxygen is produced by passing an electric current through water to which a small amount of sulphuric acid has been added. The process is known as the electrolysis of water (see Fig. 7-4). The electrical energy must be supplied from a battery or other source of direct current, and the oxygen will collect over the positive electrode (known as the anode) while the other component of water, hydrogen, will collect over the negative electrode (known as the cathode).
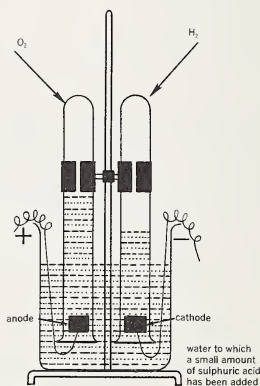

Fig. 7-4. Oxygen may be prepared by decomposing slightly acidified water by electricity.

PROPERTIES OF OXYGEN

When we discuss the characteristics of any gas, we usually consider its physical properties such as its color, its odor, its weight as compared with an equal volume of air, the conditions under which it may be changed to a liquid and a solid, and its solubility in water. In addition, we consider its chemical properties which include its ability to combine with other elements or compounds.

Physical Properties of Oxygen -- Oxygen is

a colorless, odorless gas, a little heavier than air. It is slightly soluble in water. Because of this solubility, fish are able to obtain oxygen from water and thus "breathe" by means of the gills which they possess. Because it is only slightly soluble, we can collect the oxygen produced in the laboratory by water displacement.

Oxygen is difficult to liquefy because it requires a temperature of $-183°C$ before the colorless gas changes its physical state to a pale blue liquid. If the temperature is further reduced to $-219°C$, it becomes a bluish white solid.

Chemical Properties of Oxygen – Oxygen is considered a very active element because it forms compounds with most other elements. The compounds are known as oxides. Many of these reactions take place at room temperature, some slowly and some rapidly. In all cases, the reaction will be much faster in an atmosphere of pure oxygen than in air.

Oxygen reacts with many elements known as metals according to the following equations:

$$2Na + O_2 \rightarrow Na_2O_2$$
$$2Mg + O_2 \rightarrow 2MgO$$
$$4Al + 3O_2 \rightarrow 2Al_2O_3$$
$$2Zn + O_2 \rightarrow 2ZnO$$
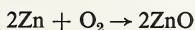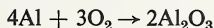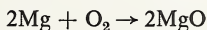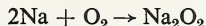
Oxygen also reacts with elements known as nonmetals according to the following equations:

$$S + O_2 \rightarrow SO_2 \uparrow$$
$$C + O_2 \rightarrow CO_2 \uparrow$$
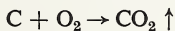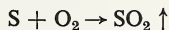
If there is not sufficient oxygen to give a complete reaction in the case of carbon, the reaction will be:

$$2C + O_2 \rightarrow 2CO \text{ (carbon monoxide)}$$
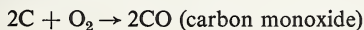
USES OF OXYGEN

Oxygen has been described as the gas of a thousand uses. While this might be an exaggeration, it is merely meant to indicate that oxygen is one of the most important gases in industry and life itself.

Fig. 7-5. An oxyacetylene torch may be used to cut through steel. The temperature of the flame is usually 3000°C to 3500°C.

Through research and experimentation since oxygen was first discovered, many uses have been found for the gas in its pure state. The two great uses, however, are: (1) in industry where the gas is used in conjunction with acetylene gas for the welding and cutting of steel and other metals and (2) in the field of medicine.

Actually the discoverer, Priestley, noted the possibility of these uses, for he observed that material that was combustible in air burned with increased vigor in pure oxygen. He also noted upon breathing the gas himself that breathing seemed easier and took less effort, or as he expressed it in his diary, "My breath felt light and easy."

Many modifications of these great uses have been offered to man in the last few years. Oxygen is used in industry not only to weld and cut metals but to clean scale from metals. It is also used in the process of recovering the metals from the compounds from which they are produced.

In the field of medicine, oxygen is used in the hospital in conjunction with other gases while the patient is under anesthetic. Later it is used to help him regain consciousness at

the conclusion of surgery. Oxygen administered in cases of emergency often helps patients sustain life when they have not the strength to obtain enough oxygen by breathing air. Oxygen is also administered to people suffering in cases of apparent drowning, shock, and asphyxiation from poisonous gases.

Oxygen is also used to sustain life in fliers whose planes take them beyond 10,000 feet altitude. It is also used to protect men who meet conditions in which oxygen is lacking, while performing duties in gas-flooded mines or when forced to work under water. In all these cases, oxygen in light but strong tanks is carried by the flier or the worker.

In recent years liquid oxygen explosives (known in the trade as L.O.X.) have become widely used in mining and related industries. The L.O.X. cartridge is a paper or fabric tube filled with powdered carbon. The cartridge is soaked in liquid oxygen just before it is to be used, placed in position, and then fired by a detonator. The carbon and oxygen unite to form great volumes of carbon dioxide at high pressures.

Adaptations of L.O.X. are used in mining and road building, as well as providing the propellant force for certain types of rockets.

THINGS TO REMEMBER

Oxygen makes up about one-fifth of our atmosphere. The remainder is nitrogen plus a small amount of water vapor, carbon dioxide, and the inert gases: helium, neon, argon, krypton and xenon.

The atmosphere is a mixture of gases.

Credit for the discovery of oxygen is given to Joseph Priestley. K. W. Scheele and Antoine Lavoisier contributed to the world's early knowledge of oxygen.

Oxygen may be prepared for use in the laboratory by heating potassium chlorate in the presence of a catalyst, manganese dioxide.

It may be collected by water displacement because it is only slightly soluble in water.

We may test the product by inserting a glowing splint in the gas and noting the vigorous reaction as the splint bursts into flame.

Oxygen is prepared for industry by the liquefaction of air.

Oxygen may also be prepared by the electrolysis of water to which has been added a small amount of sulphuric acid. The oxygen is collected at the anode.

Physical properties of oxygen: colorless, odorless gas, slightly soluble in water, liquefies at $-183°C$ and freezes at $-219°C$.

Chemical properties of oxygen: it is an active element and forms oxides by uniting with metals and with nonmetals.

Uses of oxygen in industry include its use in oxyacetylene welding and cutting torches, explosives, and related applications.

Uses of oxygen in medicine include treatment of patients for anesthetic shock, postoperative complications, respiratory conditions, and related ailments.

QUESTIONS

Part A

1: State what gases are found in our atmosphere and the division of the two major ones by volume.
2: What do we mean when we say oxygen is an element of vital importance?
3: Oxygen unites directly with many elements and compounds. What is the general name given to the product?
4: Describe the method used by Priestley to produce oxygen.
5: Using an encyclopedia as a source book, state at least two important contributions that are credited to Lavoisier in the field of science.
6: Using a diagram, describe how you would prepare oxygen in the laboratory using any method you choose.
7: In the preparation of oxygen from potassium chlorate we use manganese dioxide as a catalyst. What do we mean by the word "catalyst"?
8: Why is convenience of preparation rather than cost of raw material of prime importance in the preparation of a substance in the laboratory?
9: The equipment for the production of oxygen in industry is very costly. Why is the low cost of the raw material a more important consideration than the cost of the equipment?
10: What scientific principles are involved in the preparation of oxygen in industry today?
11: State the physical properties of oxygen.
12: State the chemical properties of oxygen.
13: Write equations for the reaction of oxygen with four metals and two nonmetals.
14: Write a paragraph on the uses of oxygen in industry and a second paragraph on the uses of oxygen in medicine.
15: What is the meaning of L.O.X.?

Part B

16: Suggest a method for preventing metals from oxidizing.
17: Oxygen is said to be a supporter of combustion but yet it does not burn. Using equations if you wish, explain the difference between a gas burning and one that supports combustion.
18: Wood in the form of a log is more difficult to ignite than wood in the form of shavings. Suggest a reason for this.
19: The property of oxygen that has to do with its solubility in water is valuable to certain forms of life and is also valuable to us in the laboratory. Discuss.
20: A few glowing coals may be "fanned" into flame. Explain.
21: Suggest a reason why carbon, which burns slowly in air, bursts into flame in an atmosphere of oxygen.
22: The rusting of iron is considered to be a sample of slow oxidation, while the burning of paper is considered to be rapid oxidation. State noticeable differences between the forms of combustion.
23: In question 22 would the products of the slow oxidation of iron weigh more or less than the original iron? Would the products of the combustion of the paper weigh more or less? Explain.
24: If 10 grams of magnesium were burned in oxygen, what would be the percentage increase in weight of the final product over the original magnesium?
25: What is the difference in cost of producing 100 grams of oxygen from mercuric oxide for which you pay $5.00 per 100 grams and from potassium chlorate for which you pay $3.00 per kilogram?

26: Assuming that gasoline has the formula C_8H_{18} and assuming that the products of complete combustion are carbon dioxide and water, write a balanced chemical equation representing the complete combustion in oxygen.

IN THE LABORATORY

7-6. To test the action of heat on mercuric oxide*

Into a pyrex test tube place a small amount of mercuric oxide (about the amount that can be piled on a ten-cent piece). Using a test-tube clamp, hold the base of the test tube in the flame of a Bunsen burner and heat it strongly. (a) What color change do you notice in the mercuric oxide? (b) Lower a glowing splint into the test tube and note the results. (c) Explain the results. (d) Observe the deposit on the side of the test tube and suggest what it might be. (e) Write an equation describing the reaction.

7-7. To test the action of heat on potassium chlorate in the presence of a catalyst, manganese dioxide

Arrange apparatus as indicated in Fig. 7-6. Mix about 10 grams of potassium chlorate with 5 grams of manganese dioxide, and place the mixture in the test tube. Heat gently in the flame of a Bunsen burner. Collect the gas by water displacement as indicated. Do not start collecting the gas immediately but allow the first few bubbles to pass off into the air. Why? Collect several bottles of the gas.

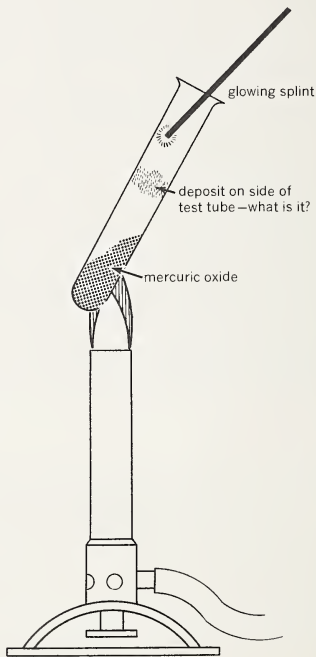*Caution: Remove the delivery tube from the water in the trough before you stop heating the test tube. Why?*

glowing splint

deposit on side of test tube—what is it?

mercuric oxide

Fig. 7-6. Testing the properties of oxygen.

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* Use fresh stock of mercuric oxide. As the oxide ages, it tends to change to a more stable form from which it is difficult to obtain oxygen by this method.

Do not attempt to inhale any of the oxygen made from mercuric oxide as the oxygen may be mixed with mercury vapor which is poisonous.

To test the properties of oxygen†

Place a small amount of sulphur on a deflagrating spoon arranged as in Fig. 7-7 and ignite the sulphur. Note the flame. Lower the burning sulphur into one of the bottles of oxygen and note the flame. Account for the difference.

Repeat the same experiment using a piece of glowing carbon instead of sulphur.

Take 3 inches of picture wire (soft iron-wire cable) and moisten one end. Dip the moist end in powdered sulphur so that as much sulphur as possible will stick to the wire. Wind the wire loosely around the deflagrating spoon, ignite the sulphur, and lower the spoon into a jar of oxygen. Note the results. Examine the oxide formed. What is the oxide called? Why did the iron oxidize rapidly in the bottle?

Caution: Use a bottle that may be discarded as it may crack during the reaction.

handle of
deflagrating spoon

cork acts as
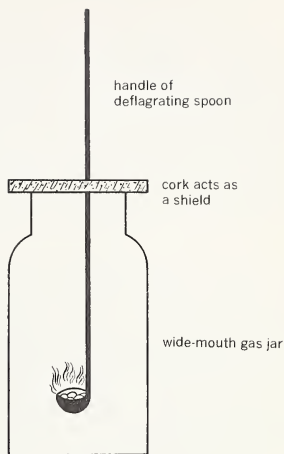a shield

wide-mouth gas jar

Fig. 7-7.

† In all these experiments be sure that the handle of the deflagrating spoon is passed through a large cork that will fit closely over the mouth of the gas bottle when the spoon is lowered into the oxygen.

CHAPTER 8 HYDROCHLORIC ACID: A TYPICAL ACID

At some time in your life you have, perhaps, heard someone described as having an acid disposition. While you may not have fully understood all that the remark implied, at least you gathered that such a description was not intended to be a compliment. The word "acid" comes to us from the Latin word *acidus* meaning "sour."

Most students would immediately question the advisability of using taste as a test for an acid, especially if they are thinking of acids in terms of concentrations such as we have in our laboratory reagent bottles. But acids are found in many other places and in solutions much more dilute than in the laboratory.

Many of our foods contain acids. Citric acid, for example, is found in lemons as well as in other citrus fruits. Tartaric acid, another organic acid, is found in grapes. The sour taste of the green apple is, in part, caused by the presence of a high percentage of oleic acid in the fruit.

Gastric juice, a digestive fluid secreted by the glands in the mucous membranes of the stomach, contains 0.3 per cent hydrochloric acid, which is necessary for the proper digestion of protein such as we find in meat, cheese, and other foods.

You are familiar with the taste of pickles which are sour because of the presence of vinegar, a dilute acetic acid.

Another property characteristic of all acids is that they react with certain complex organic substances called indicators to produce color changes. There are many such compounds, but two will be of special interest to the student of elementary chemistry. Litmus is a blue-colored solid that is slightly soluble in water. When acid is added, the color changes to a pink that is frequently described as red. Methyl orange, another organic compound, is composed of orange crystals, which dissolve readily in water. On the addition of acid the solution turns to a definite red color.

A third property of acids is their action with most metals to produce hydrogen and a salt. We have seen such a reaction in the production of hydrogen in Chapter 6. It is a common error to speak of acid attacking a metal. Possibly the reason this mistake is made is that the action is frequently observed as a result of spilling acid on a metal surface. Actually the metal is the more active element and it is the metal that forces the hydrogen from the compound.

The fourth property of all acids is their ability to react with another class of compound called a base (see Chapter 9, page 100) to form a salt and water. This reaction is called *neutralization* and is discussed in Chapter 9.

STRONG AND WEAK ACIDS

The chemist in his study of this class of compound naturally asks these questions: What is the property common to all acids? Why will all acids produce the same color changes in an indicator? What gives all acids a sour taste? Why do all acids neutralize bases? In short, what property makes an acid an acid?

If you examined the formula of several common acids such as hydrochloric acid, HCl, sulphuric acid, H_2SO_4, and nitric acid, HNO_3, you would see that all have at least

1 atom of hydrogen and therefore this must be the common property. This answer is correct to a certain degree. They do possess hydrogen, but many other compounds that are not acids also contain hydrogen.

The next step in reasoning, then, is to ask what is different about the hydrogen found in acids from the hydrogen found in other substances. The answer to this question solves our problem concerning acids.

Acids have the ability to dissociate in a water solution producing charged particles called *ions*. It is the hydrogen ion represented as H^+ that gives the acid its properties. The acids mentioned above dissociate, or ionize, according to the following:

$$HCl \rightarrow H^+ + Cl^-$$
$$H_2SO_4 \rightarrow 2H^+ + SO_4^{--}$$
$$HNO_3 \rightarrow H^+ + NO_3^-$$

Not all acids ionize to the same degree. The three acids mentioned above are considered **strong acids** because they dissociate to a high degree. By that we understand that out of every 100 molecules, practically all ionize and each molecule produces one. or more hydrogen ions as well as a nonmetallic ion.

Other acids that dissociate to a much lower degree are considered **weak acids** because they produce very few hydrogen ions. Acetic acid (CH_3COOH), sulphurous acid (H_2SO_3), and chloric acid ($HClO_3$) are examples of weak acids.

Do not confuse strength with concentration. An acid is considered strong because it ionizes freely producing a high percentage of hydrogen ions. A concentrated acid is simply one which is in solution with a small amount of water.

HYDROCHLORIC ACID

Hydrochloric acid is called a typical acid because it possesses all the properties of a typical acid mentioned above.

It is prepared in the laboratory by dissolving hydrogen chloride gas, which has the same formula —HCl, in water.

The gas is prepared by the action of concentrated sulphuric acid, H_2SO_4, on sodium chloride, NaCl.

Caution: You will observe in Fig. 8-1 that the delivery tube does not extend below the level of the water. If the tube were below the level of the water, once the reaction started the water would be forced up through the delivery tube into the reaction flask. This would happen because of the fact that hydrogen chloride gas is very soluble in water.

EXPERIMENT

8-1. *To prepare hydrogen chloride and hydrochloric acid in the laboratory*

*Caution: **This demonstration should be conducted with the apparatus in a fume cupboard with the exhaust fan in operation.***

Place about 30 grams of sodium chloride (NaCl) in a Florence flask. Add just enough water to moisten the salt. Fit the flask with a two-holed rubber stopper equipped with a long-stemmed thistle tube and a gas-delivery tube leading to a collecting vessel as shown in Fig. 8-1.

By means of the thistle tube, add sufficient concentrated sulphuric acid to cover the lower end of the tube in the flask. Avoid breathing any of the gas produced as it will unite with the moisture from the mucous membranes and produce hydrochloric acid.

*1. Collect a test tube full of the gas by the upward displacement-of-air method. Blow across the open end of the test tube and explain your observation. Caution: **Be careful that you do not inhale the gas.***

2. Half fill a gas jar with water. Direct the gas from the delivery tube so that it strikes the surface of the water for a period of a few minutes. (Note the warning above and do not let the end of the delivery tube come in contact with the water in the gas jar.)

Test the water solution of the gas with blue litmus paper and account for the change. Introduce a small piece of magnesium ribbon to the water solution of the gas and note the results. Explain.
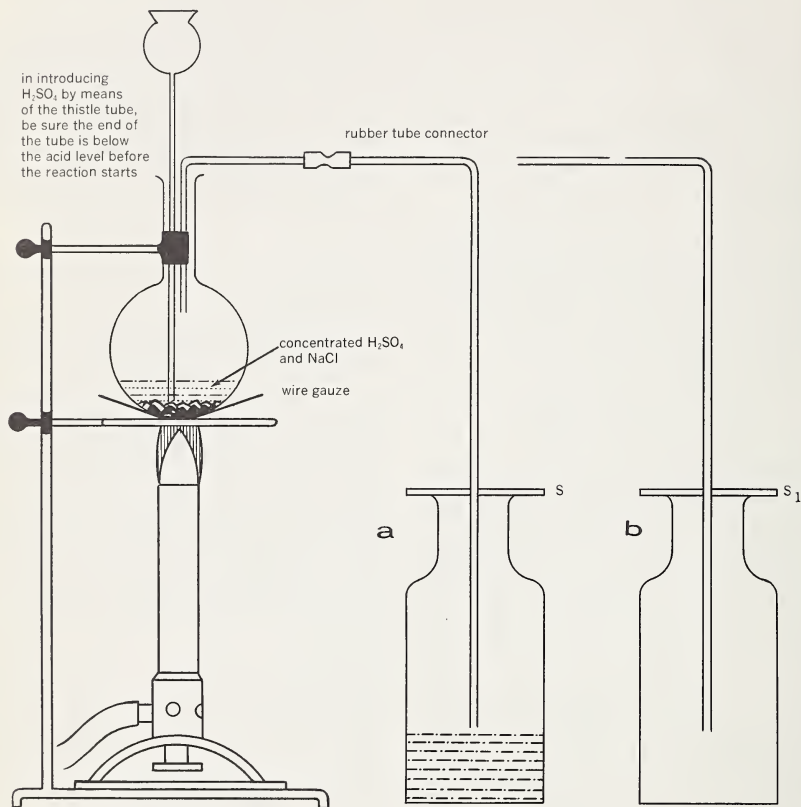
Fig. 8-1. Preparation of hydrogen chloride and hydrochloric acid in the laboratory. *Caution:* Be sure you have a wire gauze under the Florence flask. Be sure the delivery tube leading into bottle *a* does not extend to touch the water. Place a shield (*s* and *s₁*) over each of the bottles. Bottle *b* may be connected to the source by means of the rubber connector after the gas has been dissolved in the water in bottle *a*.

GENERAL PREPARATION OF ACIDS

The method outlined above for the preparation of hydrochloric acid is the method followed in preparing many other acids. In general terms it is described as treating the salt of the acid to be prepared with concentrated sulphuric acid.

The acid formed usually has a lower boiling point than sulphuric acid and can be separated from the H_2SO_4 by making use of this property.

Another method of producing certain acids is by the action of water with the soluble oxide of a nonmetal. Although hydrochloric acid is not produced by this second method, the process should be mentioned here in connection with the discussion of acids.

EXPERIMENT

To produce carbonic acid, H_2CO_3, by the action of carbon dioxide and water

Fill a beaker half full of distilled water and test it with blue litmus paper. Take a deep breath and after holding it for a moment exhale through a glass tube into the distilled water. Repeat the process several times. Test the water with another piece of blue litmus paper and note the color change.

The above reaction may be described by the equation

$$CO_2 + H_2O \rightarrow H_2CO_3$$
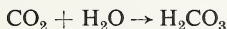
Other acids may be prepared in a similar manner, that is, by the addition of water to the oxides of the nonmetals found in the acids.

$$SO_2 + H_2O \rightarrow H_2SO_3$$
$$SO_3 + H_2O \rightarrow H_2SO_4$$
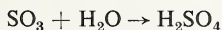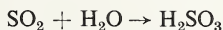
These oxides are known as *acid anhydrides.* An acid anhydride may be defined as the soluble oxide of a nonmetal that produces an acid when dissolved in water.

COMMERCIAL PREPARATION OF HYDROCHLORIC ACID*

Hydrochloric acid in industry is prepared in a method similar to the laboratory preparation, except that higher temperatures than can be withstood by laboratory glassware must be used. The higher temperature drives the reaction to completion and removes all the hydrogen from the sulphuric acid as indicated:

$$2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$$
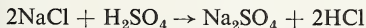
A second method used in industry for the preparation of hydrochloric acid is a direct synthesis brought about by the burning of hydrogen in an atmosphere of chlorine. Notice that in the equation, we use the molecular formula of both hydrogen and chlorine.

$$H_2 + Cl_2 \rightarrow 2HCl$$

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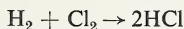
* An impure form of hydrochloric acid is known as muriatic acid. With the aid of a dictionary, account for the name.

PROPERTIES OF HYDROCHLORIC ACID

A solution of hydrogen chloride gas in water is known as hydrochloric acid because the solution shows all the general properties possessed by acids.

1. The solution has a sour taste.
2. It turns blue litmus pink.
3. It reacts with many metals to produce hydrogen.
4. Hydrochloric acid neutralizes bases such as sodium hydroxide. This will be discussed in Chapter 9.

DEFINITION OF AN ACID

By looking at the formula of hydrochloric acid, which we consider typical, we can define an acid as the union of hydrogen and a nonmetal or nonmetallic radical, which will ionize allowing the hydrogen to be replaced by a metal. This leads to the definition of an acid as being any substance containing ionizable hydrogen.

USES OF HYDROCHLORIC ACID

1. *Plating of Metals* – Many metallic articles used today consist of a cheaper metal, such as iron, covered with a thin layer of a second metal which gives the finished product a more pleasing appearance or which makes it more resistant to corrosion. For example, the container we know as a tin can is actually made of soft sheet iron which has been coated with tin. Food acids do not react with tin but form objectionable compounds when they react with iron. Chromium is used to plate automobile grilles, bumpers, and other parts of the car because of the pleasing appearance of chromium as well as its resistance to atmospheric corrosion. Galvanized iron is sheet iron coated with zinc for use in the building trades as well as for use in making articles such as pails, dippers, and other household utensils.

Before the coat of the second metal will adhere to the base metal, the base metal must be free from oxides; otherwise the plating metal does not bond successfully to the base.

One of the best ways that we know of to clean the base metal is to dip it momentarily into a bath of dilute hydrochloric acid. This process is known in the trade as *pickling* and uses great amounts of hydrochloric acid.

The action of hydrochloric acid on oxides may be quickly demonstrated by simply taking a penny that has been in use for some time and is covered with a layer of brown oxide and dipping it in a dilute solution of hydrochloric acid. You will notice that the coin loses its dull appearance after a short time in the acid.

Metals that are to be soldered should be treated in a similar fashion.

2. *Uses in Industry* – Hydrochloric acid is a valuable compound in the chemical industry as well as in the chemical laboratory. It is the starting point for many chlorides that we use, as well as being an active reagent in chemical processes. Hydrochloric acid is also used as a catalyst in certain organic reactions, such as the production of commercial glucose, which is a sugar produced from corn starch and used in the food industry. The acid also plays an important role in the production of gelatin and glue from animal skins and bones.

In addition to being an active acid, it is also a relatively cheap reagent. Such a combination as this makes it especially valuable in the field of industry.

THINGS TO REMEMBER

Hydrochloric acid is a typical acid because it has properties that are similar to the properties possessed by all acids.

The general properties of acids include:
1. A sour taste.
2. The ability to change indicators (acids turn blue litmus to pink, and methyl orange to red).
3. The ability to react with most metals to produce free hydrogen and a salt.
4. The ability to neutralize bases.

Strong acids ionize, or dissociate, to a great extent to produce ions of hydrogen written as H^+ and ions of nonmetals or nonmetal radicals which carry a negative charge.

Weak acids dissociate to a limited degree only, giving low concentrations of hydrogen ions.

Hydrochloric acid is prepared by heating a salt of the acid with concentrated sulphuric acid: $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$. This is the general method of preparing an acid.

An acidic anhydride is the soluble oxide of a nonmetal which reacts with water to produce an acid.

The commercial preparation of hydrochloric acid is similar to the laboratory preparation except it is carried out at higher temperatures and all the hydrogen is taken from the sulphuric acid: $2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$.

An acid is the union of hydrogen and a nonmetal or nonmetallic radical. It must deliver hydrogen ions in solution.

The uses of hydrochloric acid include "pickling" and cleaning of metals, acting as a reagent in the chemical and food industries, and acting as a catalyst.

QUESTIONS

Part A

1: What is the derivation and meaning of the word "acid"?
2: Name five foods that contain acid.
3: What acid is found in the gastric juices in the human body? What is the value of the acid to the system?
4: State four properties possessed by all acids.
5: What is meant by an "indicator" as used in chemistry?
6: Name two indicators and describe their appearance in acids.
7: What is meant by a "strong acid"? Give two examples.
8: What is meant when we speak of a weak acid? Give two examples.
9: What is meant when we speak of a concentrated acid?
10: What term would we use to speak of an acid that was not concentrated?
11: Why is hydrochloric acid considered a typical acid?
12: Describe how it is prepared in the laboratory.
13: What is the difference between the laboratory preparation and the preparation of the acid in industry? Give equations for each preparation.
14: What precautions do we have to observe when preparing hydrogen chloride gas in the laboratory?
15: State the general rule for preparing most acids.
16: What is an acid anhydride?
17: Name two acid anhydrides, give the formula of each, and by an equation show the reaction that takes place when each is dissolved in water.
18: State the properties of hydrochloric acid.
19: Define an "acid."
20: State the main uses of hydrochloric acid.

Part B

21: What acid with which you are familiar would it be impossible to prepare using the general method as described in this chapter?
22: Devise a method to test whether concentrated sulphuric acid is weak or strong.
23: Name five acids, give the formula of each, and show how they would ionize.
24: Certain soils found in farmlands are said to be "sour" because they are acid in nature. Devise a method by which you could test such soils to see whether or not they were acidic.

IN THE LABORATORY

To test the action of indicators in acids

Place six clean test tubes in a test-tube rack, and number them 1 to 6. Into test tubes 1 and 2 pour about 10 cc. of dilute sulphuric acid; into test tubes 3 and 4 pour 10 cc. of dilute nitric acid; and into test tubes 5 and 6 pour 10 cc. of dilute hydrochloric acid. To test tubes 1, 3, and 5 add a few drops of a blue litmus solution or place a strip of blue litmus paper in each. To test tubes 2, 4, and 6 add a few drops of a solution of methyl orange. Make a chart like the one on page 98

in your notebooks and fill in with the results observed.

If this experiment were carried out using other acids, we would find that the indicators used always gave the same results. From the above experiment what conclusion would you make concerning the color of litmus in the presence of an acid? What color is methyl orange in an acid solution? Other indicators, of course, will produce different colors. If time permits, test other indicators

	H_2SO_4	HNO_3	HCl
Litmus			
Methyl orange			

from the laboratory stock. Try this experiment with vinegar, lemon juice, and other solutions that you suspect might be acidic.

To test the action of acids on metals

In Chapter 8 the statement is made that acids react with most metals to produce hydrogen. Hydrogen is a colorless gas but it is dangerous and, if ignited when mixed with air, it will explode. *Caution: When performing this experiment use small quantities of the metals and follow directions.*

Place four clean test tubes in a rack and number them 1 to 4. Into each test tube pour about 15 cc. of dilute hydrochloric acid. To test tube number 1 add a small piece of zinc metal* about the size of the head of a wooden match. What do you observe? Hold the lighted end of a taper or splint (at least 6 inches long) over the open end of the first test tube. Be sure the level of your hand is below the rim of the test tube and that you are standing well back. What happens? The small explosion, known as a "bark," is accepted in elementary chemistry as a test for free hydrogen gas.

Repeat the experiment, placing a small piece of magnesium in tube number 2. To test tube number 3 add a small piece of aluminum and observe the results. Add a small piece of copper to test tube number 4. Make a chart like the following in your notebooks and fill in the results of your observations.

To test the ability of hydrochloric acid to clean metals

Select four coins, 2 one-cent pieces and 2 ten-cent pieces. These coins should be dull from use. Place one of the pennies in a test tube containing 20 cc. of dilute HCl. Agitate the tube so that the acid washes over the coin. Heat gently if necessary. Drain the acid and rinse the coin. Buff it with a coarse cloth and then with one with a fine weave. Compare it to the untreated one-cent piece. Account for the difference in appearance. Repeat the experiment using one of the ten-cent pieces.

|	Zinc	Magnesium	Aluminum	Copper
Visible results when metal was added to acid				
Results of lighted splint test				
Conclusions				

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* Use a commercial grade of zinc (known as "mossy zinc"). If pure zinc is used add a little copper sulphate to the solution.

# CHAPTER 9 SODIUM HYDROXIDE: A TYPICAL BASE

Bases may not be as familiar a compound to the student as acids, because they are not commonly found in nature. They form, nevertheless, an extremely important class of compounds. A base is defined as a substance containing one or more ionizable hydroxyl radicals.

THE PREPARATION OF SODIUM HYDROXIDE IN THE LABORATORY

This is carried out by the action of metallic sodium with water. Sodium, as an element, shows great activity and when it reacts with water it does so by replacing one of the atoms of hydrogen in the water.

$$2Na + 2HOH \rightarrow 2NaOH + H_2 \uparrow$$
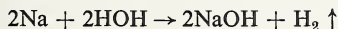
You will notice that here again we will find it much easier to follow the reaction if we write the formula of water as HOH rather than H_2O, as the OH radical remains as a unit throughout the reaction.
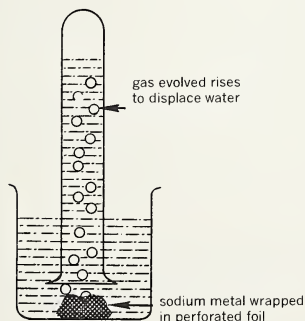
EXPERIMENT

9-1. To prepare sodium hydroxide by the action of sodium and water

Caution: **Due to the great activity of sodium, it is suggested that the teacher carry out this demonstration.**

Use a small piece of sodium not any larger than the head of a match. Once the sodium has been cut it should be replaced under kerosene in the stock bottle. It should be handled only with forceps and never touched, as a sodium burn is painful and heals very slowly.

Clamp a large inverted test tube filled with water in a partly filled water trough, in such a position that the small piece of sodium, mentioned above, wrapped in perforated lead foil may be placed under the mouth of the test tube. After the reaction between the water and the sodium is complete, test the gas that has been collected in the inverted test tube with a burning splint.
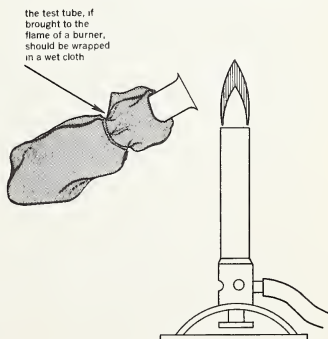

the test tube, if brought to the flame of a burner, should be wrapped in a wet cloth

gas evolved rises to displace water

sodium metal wrapped in perforated foil

Fig. 9-1. The action of metallic sodium and water.

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Take a sample of the water from the trough and test it with red litmus paper. What is the difference between the feel of a few drops of the water from the trough and pure water?

Write a skeleton equation for the reaction and balance it according to the steps set down in Chapter 6, page 71.

Another method of producing sodium hydroxide in the laboratory is to have water act on the basic anhydride. In the chapter on acids we learned that an acidic hydride was the soluble oxide of a nonmetal that united with water to produce an acid. The same general principle applies to bases, and the basic anhydride may be defined as the soluble oxide of a metal that reacts with water to form a base.

In the case of sodium hydroxide, the reaction is with the basic anhydride, sodium oxide Na_2O:

$$Na_2O + HOH \rightarrow 2NaOH$$

Other bases may be formed the same way:

$$K_2O + HOH \rightarrow 2KOH$$
$$CaO + HOH \rightarrow Ca(OH)_2$$

PROPERTIES OF SODIUM HYDROXIDE

1. A dilute solution of sodium hydroxide feels slippery to the touch. We discovered this when we tested the dilute solution of NaOH which was one of the products of reaction in the demonstration above.

2. Sodium hydroxide reacts with indicators to produce a color change. The water from the trough in the demonstration contained NaOH in solution, and we noted that it changed red litmus to blue. It would also have changed methyl orange to a yellow color.

    Another valuable indicator that we could use is phenolphthalein (pronounced fē'-nol-thal'-een), which is a white powder that dissolves quite readily in methyl alcohol. When a few drops of this solution are added to a sample of any base, a red color is produced. Phenolphthalein is colorless in acids.

3. Sodium hydroxide reacts with acids to produce a double decomposition type of reaction as in this equation:

$$NaOH + HCl \rightarrow NaCl + HOH$$
$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2HOH$$
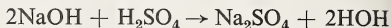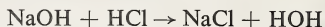
These reactions are known as neutralization.

NaOH *is a typical hydroxide*. In Chapter 8 we learned that the general properties of all acids were the result of the hydrogen ion. In the case of the bases, the slipperiness to the touch, the action on indicators, and the ability to neutralize acids is dependent upon the presence of an OH^- (hydroxyl) radical. The strength of the base may be measured by the percentage of dissociation that takes place when the compound is dissolved in water. Certain hydroxides are soluble in water, and dissociate completely producing a high concentration of hydroxyl ions. These are called strong bases. Three common strong bases are sodium hydroxide (NaOH), potassium hydroxide (KOH), and lithium hydroxide (LiOH). They ionize according to the following:

$$NaOH \rightarrow Na^+ + OH^-$$
$$KOH \rightarrow K^+ + OH^-$$
$$LiOH \rightarrow Li^+ + OH^-$$
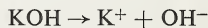
Weak bases may be soluble bases that do not dissociate to any great degree such as NH_4OH. Compounds like $Ca(OH)_2$ that are sparingly soluble may be bases, as the small part that does dissolve is sufficient to give a basic reaction.

Sodium hydroxide is considered a typical base because it ionizes to produce a high concentration of hydroxyl ions and, therefore, to a very marked degree exhibits the properties common to all bases.

NEUTRALIZATION

The double replacement reaction that takes place when a base and an acid react is known as neutralization because the characteristic part of the acid, the hydrogen ion,

and the characteristic part of the base, the hydroxyl ion, unite to form a neutral compound, water, that does not dissociate to any marked degree.

The point at which neutralization is complete when such a reaction is being carried out is recognized when the indicator litmus begins to change color. In conducting a demonstration of neutralization, we should use the indicator litmus because it changes color at a point closer to a true neutral solution than does methyl orange or phenolphthalein. Methyl orange makes its color change while the solution is slightly acidic, while phenolphthalein makes its change when the solution is slightly basic. Both these indicators have an important place in certain chemical processes because of the properties which they possess, and both are useful to us for rough work.

EXPERIMENT

To show the neutralization of an acid by a base

Prepare a solution of hydrochloric acid by taking 17 cc. of concentrated HCl and adding sufficient distilled water to make 65 cc. of solution. Prepare a solution of sodium hydroxide by weighing out 6 grams of NaOH and dissolving it in 45 cc. of water. When the sodium hydroxide is completely dissolved, add sufficient water to make the total volume of the solution 65 cc.

Pour 15 cc. of the acid into a clean beaker and add a few drops of a solution of blue litmus. Note the color change.

Now pour about 12 cc. of the NaOH solution into the acid and litmus solution. What do you notice about the color where the base solution strikes the acid solution?

Stir with a clean glass rod.

Now add the base solution very slowly (an eye-dropper will allow you to add it a few drops at a time) stirring the mixture constantly. Stop at the moment that the acid solution's color shows the first sign of retaining a purplish color. If you go past that point, add a few drops of the acid solution to bring the color back to a point between red and blue. Be sure you use a clean dropper for the

acid solution. When this color is obtained, we may assume that the solution is neutral.

Place the contents on an evaporating dish and, using a low flame, evaporate to dryness. Using a moistened finger to obtain a sample, taste the residue. Describe the taste.

The reaction may be described by the equation:

$$NaOH + HCl \rightarrow NaCl + HOH$$
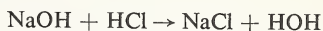
The compound, other than water, produced by the neutralization reaction is called a salt. It will be discussed further in Chapter 10. The main fact to remember concerning neutralization at this point is that it is the action of an acid and a base that produces water. This water is neutral because it is composed of the active part of an acid and the active part of a base forming a compound that does not ionize or dissociate to any degree. The salt depends on the acid and on the base used.

The definition of **neutralization may be stated as the action of an acid and a base to produce a salt and water.**

THE PREPARATION OF NaOH IN INDUSTRY

Sodium hydroxide, also known as caustic soda, is an important compound used in many of our industrial processes. It is produced commercially by passing an electric current through a solution of brine. This is known as the electrolysis of brine. Brine is a concentrated solution of sodium chloride. While the mechanics of the process are simple, the theory involved is beyond the level of this course. The reaction, however, may be expressed in a single equation:

$$2NaCl + 2H_2O \rightarrow 2NaOH + H_2 + Cl_2$$
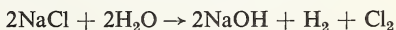
The sodium hydroxide which is the substance that the chemist wished to produce is known as the primary product. There are also two secondary products, hydrogen and chlorine. Both of these have value to industry. They are referred to as by-products.

The sodium hydroxide produced in industry, like that produced in the laboratory, is in solution in water. When the water is evaporated the NaOH remains as a white crystalline substance. If it is not stored in an air-tight container, it exhibits the property known as deliquescence. That is, it absorbs sufficient moisture from the air to put itself into a concentrated solution.

This may be demonstrated by taking a small piece of sodium hydroxide from the reagent bottle and placing it on a watch crystal. Shortly after the solid compound has been exposed to the air, you will notice that the outside surface has become coated with moisture. After several hours of exposure, there will be a small pool of the concentrated sodium hydroxide solution surrounding the original sample.

*Caution: **Do not touch the sodium hydroxide. Use forceps to manipulate it at all times. Sodium hydroxide is caustic, that is, it burns the skin and, like the burn from the element sodium, it is very painful and slow to heal.***
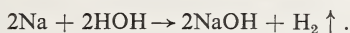
USES OF SODIUM HYDROXIDE

Sodium hydroxide is one of the world's most useful compounds. Industry on this continent requires nearly two million tons of sodium hydroxide each year. The following list includes some of the more important uses.

1. About one-fifth of the production of sodium hydroxide is used in the manufacture of rayon and cellulose film. Rayon is a synthetic, or man-made, fiber used to make a thread that is woven to produce a fabric. While rayon is similar in many respects to cotton, it was first introduced as "artificial silk."

   Cellulose film, of which cellophane is one form, is similar in its composition to rayon. The main difference in the process of manufacture occurs during the last step which results in the production of a sheet rather than a thread.

2. The chemical industry uses great quantities of caustic soda. Sodium hydroxide itself is not expensive and it is used as a starting point in the manufacture of many sodium compounds that, among others, are produced by this industry.

3. Before the days of detergents, most of our cleansing agents were in the form of some type of soap. Great amounts of sodium hydroxide are still used in the manufacture of this product.

   Lye, a common household cleanser which is used to clear drains that have become clogged with greasy substances, contains about 94 per cent sodium hydroxide. Lye, because of its high percentage of caustic soda, produces painful burns if it comes in contact with the skin. It should be kept out of the reach of children.

4. At one stage in the refining of gasoline, sulphuric acid must be added to remove certain impurities. It is obvious that no trace of the sulphuric acid can be left in the finished product. To ensure this, measured quantities of caustic soda are used to neutralize any remaining acid.

5. Wood pulp, being prepared for the manufacture of paper, is "digested" or treated with sodium hydroxide. In Canada, the paper industry is an important one and it uses much of the sodium hydroxide produced in this country.

6. Sodium hydroxide is used in many instances in the manufacture of fine cotton textiles. The fabric or the fibers used in its manufacture are treated with sodium hydroxide and then dried under tension. This gives the finished cloth greater strength, makes it more receptive to dye, and also gives it a desirable sheen that it would otherwise lack. The process is known as mercerizing and was developed by John Mercer, an English chemist who lived during the nineteenth century.

THINGS TO REMEMBER

Sodium hydroxide is prepared in the laboratory by the action of sodium and water:

$2Na + 2HOH \rightarrow 2NaOH + H_2 \uparrow$.

The properties of sodium hydroxide are:
1. In a dilute solution it feels slippery to the touch.
2. It reacts with indicators to turn red litmus blue; red methyl, orange to yellow; and colorless phenolphthalein to red.
3. It neutralizes acids.

Sodium hydroxide is considered a typical base because it has the above properties which are found common to all bases.

Neutralization may be defined as the action of an acid and a base to deliver a salt and water. The water is neutral because it is composed of the hydrogen ion, which is the active part of an acid, and the hydroxyl ion, which is the characteristic part of a base. They form a compound, water, which ionizes to a very limited degree only.

A base may be defined as a substance that can supply hydroxyl ions.
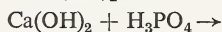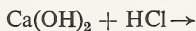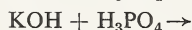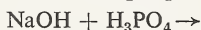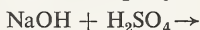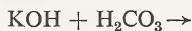
Much sodium hydroxide is used in the chemical industry; in the manufacture of rayon, cellulose films, and other related products; in the textile industry; in the manufacturing of soap; in the refining of gasoline; and in the preparation of wood pulp.

QUESTIONS

Part A

1: Describe the preparation of sodium hydroxide in the laboratory and give the balanced equation to illustrate your answer.
2: How would you test the results to show that a base had been produced?
3: Describe a second method of preparing bases in the laboratory and give equations to illustrate the method for three bases.
4: State three general properties of a base.
5: Why is sodium hydroxide said to be a typical base?
6: Explain what is meant by the process known as "neutralization."
7: Why is water considered neutral?
8: Why did we use the indicator litmus to show neutralization rather than the indicator phenolphthalein?
9: Write a paragraph on the preparation of NaOH in industry. Include a discussion of the method, the product, and the by-product.
10: What is meant when we say a substance is deliquescent?
11: What precautions must be taken when storing a deliquescent substance?
12: State precautions that must be taken when handling sodium hydroxide.
13: Discuss the use of sodium hydroxide in the chemical industry.
14: Name some commercial products that are mainly sodium hydroxide.
15: What use is made of sodium hydroxide in the textile industry?
16: Explain what is meant by "mercerizing."
17: Discuss the part played by sodium hydroxide in the business of keeping the nation clean.
18: How is sodium hydroxide used in refining gasoline? Illustrate with equations.
19: How is sodium hydroxide used in the manufacture of paper?

Part B

20: Give reasons that would lead to the statement that sodium hydroxide is not an expensive chemical reagent.

21: If you examine a stock bottle that contains free sodium you will notice that the metal is stored under kerosene. Suggest a reason for this method of storage.

22: Why is metallic sodium not found in natural deposits?

23: Why do we use sodium chloride as a starting point for most sodium compounds?

24: If you spilled an acid on your hand, it would be dangerous to attempt to neutralize it with sodium hydroxide. Why? What treatment would you give to the burn?

25: If you were given several white solids that were either acids or bases how would you distinguish between the acids and the bases?

26: Do we consider the oxides of all nonmetals as basic anhydrides?

27: Complete and balance the following equations:

$$KOH + H_2CO_3 \rightarrow$$
$$NaOH + H_2SO_4 \rightarrow$$
$$NaOH + H_3PO_4 \rightarrow$$
$$KOH + H_3PO_4 \rightarrow$$
$$Ca(OH)_2 + HCl \rightarrow$$
$$Ca(OH)_2 + H_3PO_4 \rightarrow$$

28: How many grams of sodium hydroxide would be formed by the action of 92 grams of metallic sodium with an excess of water? How many grams of hydrogen would be produced during the reaction?

29: What weight of calcium hydroxide would be required to neutralize 147 grams of sulphuric acid?

30: Supposing that during the refining process of gasoline it was found that the gasoline contained 5 grams of sulphuric acid in each gallon, what weight of caustic (sodium hydroxide) would be required to neutralize the acid present in 500 gallons of the gasoline?

IN THE LABORATORY

To test the actions of bases on indicators

Place six clean test tubes in a rack and number them 1 to 6. Pour 10 cc. of a dilute solution of sodium hydroxide into each of test tubes 1, 2, and 3 and about 10 cc. of a solution of calcium hydroxide into each of test tubes 4, 5, and 6. To test tubes 1 and 4 add a strip of red litmus paper and note the result. To test tubes 2 and 5 add a few drops of methyl orange indicator and note the results. To test tubes 3 and 6 add a few drops of a solution of phenolphthalein. Phenolphthalein, a white powder, dissolves readily in methyl alcohol and may be used in this form as an indicator. Note the results. Make a chart like the accompanying one and fill in with the colors observed. Where color changes took place, mention them.

Keep the above solutions for the next experiment.

|	NaOH	$Ca(OH)_2$
Litmus		
Methyl orange (acidified)		
Phenolphthalein		

To demonstrate the neutralization of a base by the action of an acid

To each of the six test tubes used above which contain the basic solutions, add slowly sufficient dilute hydrochloric acid to cause the indicators to change color. (The phenolphthalein will become colorless or slightly milky in appearance.) From your knowledge of Chapter 8, what is the state of the solutions now?

Using an eye-dropper, add to test tubes 1, 2, and 3 sufficient sodium hydroxide solution to cause a slight change in the color of the indicator but not sufficient to become the color that would indicate a basic solution. Do the same with test tubes 4, 5, and 6, using a solution of calcium hydroxide. What would you say was the state of the solutions in the test tubes at this stage?

Describe what has taken place by completing the following equations:

$NaOH + HCl \rightarrow$

$Ca(OH)_2 + HCl \rightarrow$
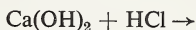
Define neutralization.

CHAPTER 10 SODIUM CHLORIDE: A TYPICAL SALT

By now you have learned that one of the products of neutralization is called a salt. Possibly you have thought that there was only one salt: common salt or table salt, sodium chloride. Actually there are hundreds of salts, and sodium chloride is simply the most common. Other examples of salts are sodium sulphate Na_2SO_4, potassium bromide KBr, ammonium carbonate $(NH_4)_2CO_3$, and sodium bicarbonate $NaHCO_3$.

In this chapter we will discuss a few facts concerning sodium chloride because it exhibits properties common to most salts.

From the examples above, you can see that *a salt may be defined as a compound formed by the union of a metal or metallic radical and a nonmetal or a nonmetallic radical.* The non-metallic radicals, SO_4^{--}, CO_3^{--}, NO_3^{-}, and others, are sometimes referred to as acid radicals because they are found in chemical combination with ionizable hydrogen, thus forming acids.

The salt, sodium chloride, has been known since antiquity. To the best of our knowledge it has always been used in the preparation of foods and in the preserving of foods. In addition, early in man's history he used common salt in the tanning of leather and in the making of pottery. Today it is considered by the chemist and the industrialist alike as the most important compound in its class.

OCCURRENCE

Great deposits of rock salt, known as *halite*, are found throughout the world. Canada has valuable deposits in New Brunswick and Ontario as well as in Northern Alberta and Saskatchewan. From these deposits comes most of the salt used in this country. The sea itself is an inexhaustible supply of salt, as by far the greatest percentage of soluble substance in it is sodium chloride.

COMMERCIAL PRODUCTION

Salt used in industry may be produced by one of several methods.

1. In some localities salt is mined in a method similar to that used in the mining of coal. It is brought out in blocks and chunks which are crushed and refined (Fig. 10-1).
2. In other operations a well is drilled to the deposit, water is forced down to dissolve the salt, and it is pumped out as a concentrated solution of brine.
3. In some areas on the coast sea water is collected in long shallow troughs and the salt is removed by evaporation.

In none of these operations is the salt produced chemically pure and a refining process must follow.

THE REFINING PROCESS

One of the great advantages of bringing salt to the surface in the form of brine is that many of the insoluble impurities of the salt deposit are left below.

The brine, which contains 26 per cent sodium chloride, is roughly six times as salty as the ocean. It is piped to large tanks where it is treated chemically to remove most of the remaining impurities. The purified brine is then pumped to large vacuum pans which are considered to be the heart of the refining

Fig. 10-1. In an underground salt mine chunks of salt are blasted from the deposit by dynamite. The chunks are loaded in mine cars and transported to the processing plant.

process. In these large tanks or kettles water is boiled off under reduced pressure. Because water will boil, that is, change rapidly into a vapor state at a temperature much lower than 100°C when the pressure is reduced, this method of treatment greatly reduces the cost of production of the finished product.

As the water passes off from the pans, the concentration of the salt in the remaining solution increases to such an extent that much of the sodium chloride precipitates out. The final product of the pans is known as "slurry." The slurry is further purified, dried, and screened into the various grades of salt that appear on our markets.

Sodium chloride, with even a trace of some of the other salts frequently found with halite, especially the chlorides of magnesium or calcium, tends to cake in damp weather. These impurities must be removed so that the salt will pour freely in a humid atmosphere.

Most table salt sold in Canada and used for the seasoning of food contains, as an additive, a fractional per cent of potassium iodide (KI).* This addition to our diet ensures a sufficient supply of the element iodine that is required by the thyroid gland to manufacture thyroxin. A deficiency of this hormone, thyroxin, leads to the condition known as goiter. Because of the presence of the iodide the package is labeled "iodized salt."

While salt is not classed as a food, sodium chloride is necessary in the diet. From it the body manufactures the hydrochloric acid required for the successful digestion of protein. Also, unless the salt balance in the blood is correct, the red corpuscles fail to supply the life-giving oxygen to the tissues and remove the waste materials.

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* Notice that the additive is potassium iodide, a compound, and not the element iodine. The iodide compound is a source of iodine for the body

As a great deal of salt is lost from the body when a person perspires freely and as it is important to maintain a proper balance, the armed forces, many factories, and some schools distribute salt tablets to their members during the summer season.

SODIUM CHLORIDE IN INDUSTRY

Because of the stress put on the uses of salt in the preserving and seasoning of food, one is likely to ignore the uses of sodium chloride in industry. Only 10 per cent of the salt produced is used in connection with food.

The chemical industry uses much salt as a starting point for the manufacture of many of the compounds of sodium that play such an important role in our daily lives. Sodium hydroxide; sodium carbonate, known as washing soda; sodium bicarbonate, known as baking soda; hydrochloric acid; and many other valuable compounds all use NaCl as the starting point in their production.

Salt is used in the manufacture of drugs, glazes for pottery, synthetic rubber, dyes, fertilizers, and other materials that we find valuable in our daily living.

SALT AS A CLASS

The chemist, of course, is mainly interested in common salt for itself and for the part it plays in life and in industry. But he is also interested in the hundreds of other members of this great class of compounds, their preparation and their properties. Some of these other salts are found in nature to the same extent as sodium chloride.

There are six common methods used by the chemist to prepare a salt in the laboratory as well as in industry. In the list that follows, the example given is not the only method for the preparation of that particular salt. The equations serve only as an illustration of a possible method, and the salt is underlined in each case.
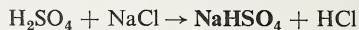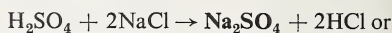
A salt may be produced by:

1. The action of an acid and a base producing a salt and water (review Chapter 8, page 92).
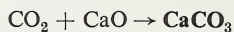
$$HCl + NaOH \rightarrow \underline{\mathbf{NaCl}} + HOH$$

$$H_3PO_4 + 3KOH \rightarrow \underline{\mathbf{K_3PO_4}} + 3HOH$$
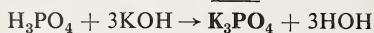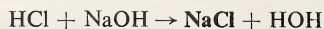
2. By the action of a metal and a nonmetal (review Chapter 3, page 33).

$$Fe + S \rightarrow \underline{\mathbf{FeS}}$$
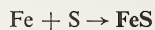
3. The action of concentrated acid on the salt of a more volatile acid (review Chapter 8, page 93).

$$H_2SO_4 + 2NaCl \rightarrow \underline{\mathbf{Na_2SO_4}} + 2HCl \text{ or}$$

$$H_2SO_4 + NaCl \rightarrow \underline{\mathbf{NaHSO_4}} + HCl$$

4. The action of a basic anhydride and an acid anhydride.

$$CO_2 + CaO \rightarrow \underline{\mathbf{CaCO_3}}$$

You will note that this reaction is closely related to the process of neutralization (review Chapter 9, page 101). The main difference appears to be that the reaction lacks water both on the left side of the equation as well as on the right side. Actually, while this is not shown in the equation, water vapor from the air enters into this reaction.

5. The reaction of an acid with the oxide of a metal (see the experiment following this list).

$$2HCl + CuO \rightarrow \underline{\mathbf{CuCl_2}} + H_2O$$
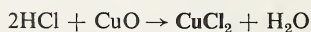
6. The reaction of two soluble salts which exchange their nonmetals in a reaction of the double decomposition type (review Chapter 2, page 22).

$$AgNO_3 + NaCl \rightarrow NaNO_3 + AgCl \downarrow .$$
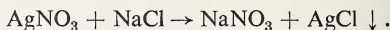
The NaNO_3, being soluble, is said to remain "up" in solution, while the AgCl, being insoluble, is said to come "down" as a precipitate.

EXPERIMENT

To show the production of a salt from the action of an acid and an oxide

In this demonstration we shall use a known solution for comparison. By that we mean that we shall make a solution of known concentration of the salt that we expect to produce in

the reaction between the oxide and the acid. This solution will be kept as a standard of comparison.

1. *To set up the known solution take 2 grams of anhydrous (dry) cupric chloride ($CuCl_2$) and dissolve it in 15 cc. of water in a test tube. Set this aside until later. Mark it as test tube A.*
2. *Take 1.10 grams of cupric oxide (CuO) and place it in a test tube the same size as test tube A.*
3. *Add hydrochloric acid slowly until the cupric oxide is completely dissolved, then add water until the level of the solution in the second test tube (test tube B) is the same as it is in test tube A.*
4. *Compare the appearance and color of the contents of the two tubes.*
5. *Clean two iron nails by dipping them in hydrochloric acid. Place one of the nails for a few moments in test tube A and the other nail in test tube B. Be sure that you keep the nails the same length of time in the solutions.*
6. *Remove the nails and compare the appearance. Can you suggest what is deposited on each nail?*
7. *Filter off any residue that appears in the bottom of the test tubes or in suspension in the solutions. Keep the solutions separate, still marked A and B.*
8. *Add silver nitrate to each of the solutions until the precipitation is complete. Decant (pour off) the excess liquid and wash the precipitates. Compare the appearance of the two precipitates.*
9. *Place the washed precipitates in direct sunlight for a few moments and observe results.*

Because the contents of test tube A, known to contain a solution of cupric chloride, reacted in a fashion similar to that of test tube B, what conclusion should we arrive at concerning the original reaction of the cupric oxide and hydrochloric acid?

Many of the experiments indicated above that were carried out in previous chapters were originally introduced to illustrate some point other than the production of a salt. From this you should learn that much information may be gained from a single reaction, and make a practice of looking for more than the obvious.

THINGS TO REMEMBER

Sodium chloride, common salt, is merely one compound of a great class known as salts.

A salt is composed of the union of a metal and nonmetal or a positive radical and a nonmetallic radical.

Sodium chloride is obtained from natural deposits of halite in New Brunswick, Ontario, Northern Alberta, and Saskatchewan.

In modern commercial production the halite is either mined by being brought to the surface in blocks or it is dissolved and pumped to the surface as brine.

The halite contains certain impurities that must be removed before the salt may be used.

The body requires salt in order that the normal functions of the blood and digestive juices may be carried on.

Industry uses great amounts of salt both as a source of sodium and chlorine and in the production of the various compounds of those elements. Salt is also used at various stages in the manufacture of synthetic rubber, drugs, dyes, fertilizers, glazes for pottery, and other things.

There are six common methods used in the laboratory for the producing of salts:

1. Neutralization.
2. Action of a metal and nonmetal.
3. Action of concentrated sulphuric acid on a salt.
4. Action of basic and acidic anhydrides.
5. Action of an acid and the oxide of a metal.
6. The reaction of two soluble salts.

QUESTIONS

Part A

1: Sodium chloride or common salt has been known since ancient times. List some of the uses of salt today that have come to us from the peoples of bygone eras. What uses do we make of salt that the ancients did not employ?
2: What areas in Canada supply us with our salt?
3: What is halite? How does it differ from table salt?
4: Describe two methods used to obtain halite.
5: Outline the refining process of table salt from the obtaining of halite to the packaged product.
6: What use does the body make of sodium chloride?
7: List some of the uses of sodium chloride in industry today.
8: Using equations to illustrate your answer, state six processes that we may use in the laboratory to obtain a salt.

Part B

9: From your observation of the substance and its reactions, list five properties of sodium chloride.
10: Describe a process that will illustrate the production of a salt from the action of an oxide of a metal and an acid.
11: Write the equation and name the salt formed when:
  a. Hydrochloric acid reacts with zinc.
  b. Sulphuric acid reacts with sodium chloride.
  c. Nitric acid reacts with sodium hydroxide.
  d. Calcium hydroxide reacts with phosphoric acid.
12: If a solution containing 50 grams of pure sodium hydroxide is neutralized by the action of sulphuric acid, what weight of sodium sulphate would be formed?
13: How many grams of sodium chloride would be formed by the action of 75 grams of sodium hydroxide with an excess of hydrochloric acid?
14: If a solution of hydrochloric acid was only 10 per cent HCl by weight, how many grams of such a solution would be required to produce 100 lb. of pure sodium chloride from sodium hydroxide?
15: What weight of NaCl would be in 5 kilograms of sea water which contains 2.45 per cent sodium chloride by weight?
16: What weight of silver nitrate would be required to precipitate the chloride in question 15?
17: What weight of silver chloride would be produced by the reaction in question 16?

IN THE LABORATORY

To produce a salt from the action of an acid and an oxide of a metal

This was demonstrated by your teacher during the presentation of this chapter. Follow the directions given for this experiment on page 108 of this text and conduct the experiment for yourself.

To test for a chloride

This test is based on certain properties of silver chloride.

1. Silver chloride is insoluble in a water solution.
2. Silver chloride is soluble in ammonium hydroxide.
3. Silver chloride is insoluble in nitric acid.
4. Silver chloride darkens rapidly in sunlight.

As one step in establishing a test, take a known chloride and treat it in the manner described below noting the results.

Take a small amount of sodium chloride (the amount you can pile on a 25 cent piece) and dissolve it in half a test tube of water. Add several drops of a solution of silver nitrate and note the results. Continue to add silver nitrate until no further precipitate is formed. Write the equation for the reaction (it is a double decomposition).

Divide the contents of the test tube into three equal parts.

To the first part add a solution of ammonium hydroxide. Shake the tube and note the results.

To the second tube add 10 or 15 cc. of dilute nitric acid. Note the results.

Set the third tube in direct sunlight for a few minutes and note any changes.

By trying this process with other chlorides, chemists have found that the results are characteristic of all chlorides and the process is accepted as a test for chlorides.
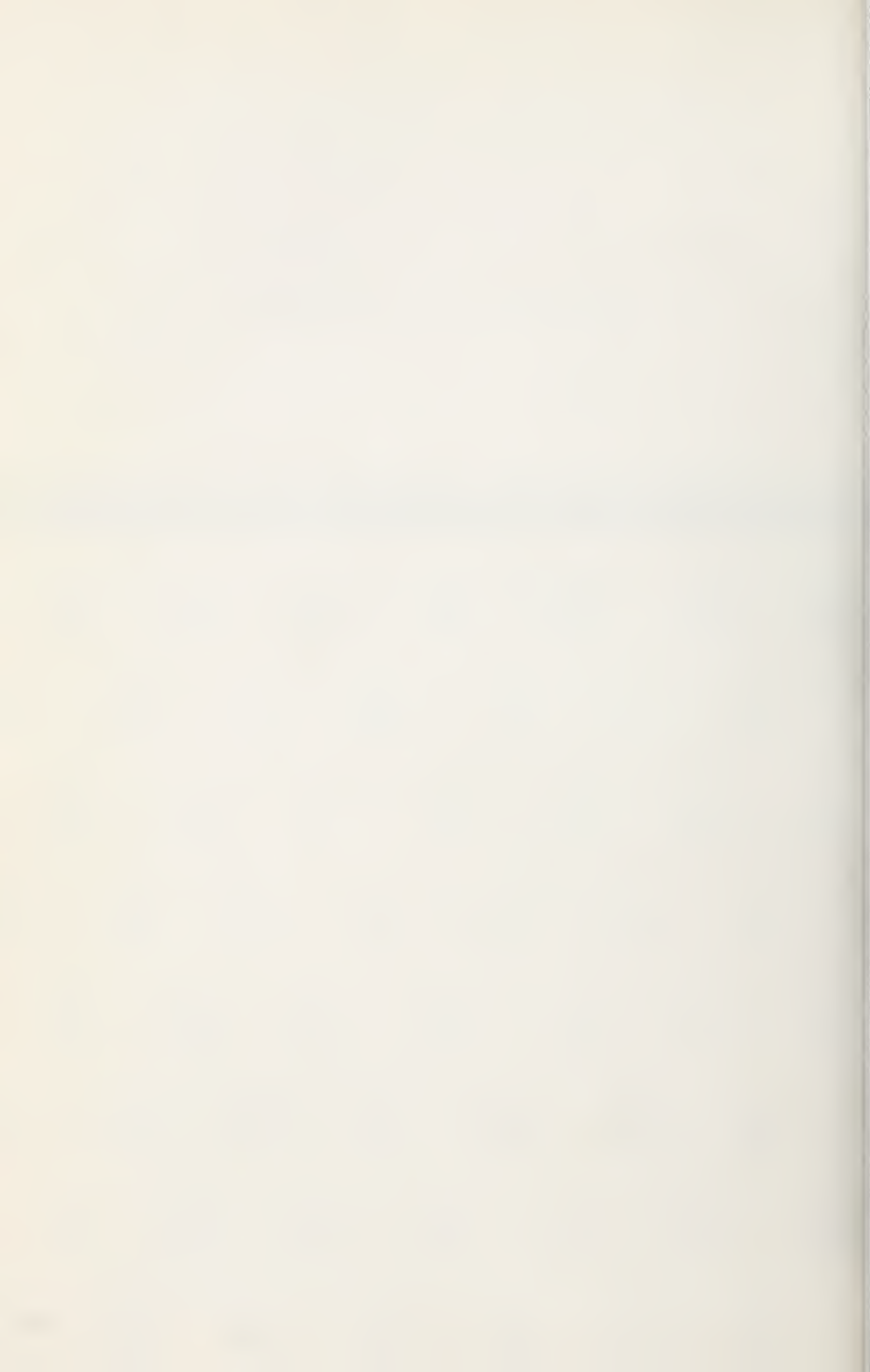
Have your teacher give you samples of salts which may or may not be chlorides. Test them following the procedure described above and have your results checked by your instructor.

An Introduction to Physics

Mechanics may be defined as a study of the effects of forces on matter, and of motion. From this definition we can see what a vast area is encompassed by the subject.

Since mechanics deals with forces, it is only proper to begin with molecular forces and introduce some of the phenomena based on molecular attraction. Terms such as *boiling*, *osmosis*, *cohesion*, *adhesion*, *surface tension*, and *capillarity* are explained in terms of molecular attraction.

In this unit, we shall limit our studies to pressures in liquids at rest and pressures in gases.

However, it should be noted that fluids in motion do not behave in the same way as fluids at rest.

CHAPTER **11** ## MOLECULAR FORCES IN LIQUIDS

The realm of the invisible, dealing with particles too small to be seen even through the most powerful microscope, is not a new concept. Since the peak of the Grecian civilization, thinkers have pondered about the nature of atoms, the tiny particles which compose matter. Democritus in the fifth century wrote:

> Atoms are infinite in number and infinitely varied in form. They strike together and their lateral motions and whirlings are the beginning of worlds.
> The varieties of all things depend upon the varieties of their number, size and aggregation.

It should be noticed that the Greeks loved to theorize, but did little experimenting in the sense that we do today. Some of their theories, however, have now been proved experimentally and are accepted by ourselves.

For 2200 years after Democritus, the view that matter is composed of particles remained a hypothesis, supported by some and opposed by others. In the early part of the nineteenth century, scientists such as John Dalton, Amadeo Avogadro, and others secured experimental evidence strongly favoring the atomic theory. This evidence accumulated throughout the century, and today we do not doubt that atoms are real. We define an atom as the smallest particle of an element that can take part in a chemical change.

Today we are aware of the fact that there are over one hundred different kinds of atoms which combine to form hundreds of thousands of different kinds of *molecules*. These molecules cannot be broken up without changing the nature of the substance. *A molecule is the smallest particle of any pure substance that still has the properties of the substance.*

SIZES OF MOLECULES

Although molecules are so small that we can never hope to see one even through the most powerful microscope, yet their diameters may be estimated by indirect methods. It is difficult to imagine how minute these values are. The distance between the centers of the neighboring molecules in water is about one hundred-millionth of a centimeter (more conveniently written 10^{-8} cm.). Actually, each molecule is so small that 100 million molecules placed side by side would cover a distance of only a half-inch.

MOLECULAR MOTION

We have evidence that molecules making up matter are always in motion. We know that it is possible to compress a large volume of a gas into a small container. This leads us to believe that the gases consume vast areas of space and consist of widely scattered molecules. These molecules must be in very rapid motion because they exert a pressure against the walls of a container.

If we turn the valve of a gas stove the odor of the gas is soon detected in all parts of a closed room. This diffusion is further evidence that the molecules of a gas are not tightly joined together but, like a swarm of flies, are relatively free to travel and dart about on zigzag paths, making impacts with

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one another and with the walls of the container. Because gases expand indefinitely, they have neither fixed volume nor shape.

Solids have definite volumes and definite shapes because their molecules are closely packed together. The molecules cannot move about freely. Their motion is generally restricted to vibrating back and forth about a fixed position.

Liquids have definite volumes but indefinite shapes. A gallon of gasoline occupies the same volume whether it be contained in a cylindrical tank or a rectangular one. Since a liquid conforms to the shape of the container, we say that it is not rigid.

If we drop a lump of sugar into a glass of water, the sugar will soon disappear, and by tasting the liquid we shall find that sugar is distributed throughout the water. This shows that the molecules of water must be relatively farther apart than the molecules in solids, to enable the molecules of the sugar to move about freely between them.

Not only are the molecules farther apart, but they are also in motion. If you leave water in a shallow pan, the water soon disappears. It has left the pan in particles too small to be seen. Because the particles, or molecules, left the pan, they must have been in motion.

THE BROWNIAN MOVEMENT

Perhaps the most convincing proof of molecular motion in liquids was discovered in 1827 by an English botanist, Robert Brown (1773–1858). He placed some pollen grains in water and dropped a bit of this suspension on a small glass slide. When he looked at the material under a microscope, he found that these small particles were in rapid motion. They vibrated in a haphazard way. The path of one particle would be like that shown in Fig. 11-1. Many scientists repeated the experiments using different solids and different liquids. It was found in these experiments that the smaller the particles, the faster they moved. The rate of motion did not depend on the kind of particles, but only on their weight. It was also found that heating a substance increases the speed of the molecules. This haphazard movement of the particles is caused by collisions with
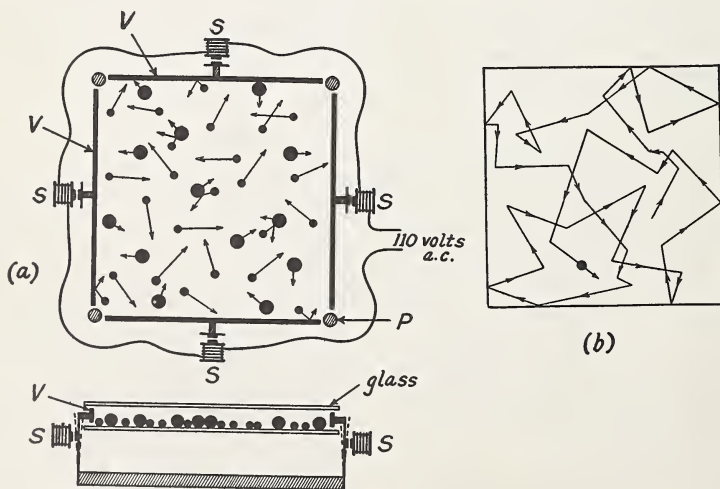

Fig. 11-1. Mechanical model illustrating the haphazard path of a single large particle caused by its collisions with the molecules of the liquid in which it is suspended.

rapidly moving molecules of the liquid in which they are suspended, and is known as the **Brownian movement.**

With all the evidence collected throughout years of experimentation, scientists have agreed that *all matter is made up of molecules and that the molecules are always in rapid motion.*

EVAPORATION

There is plenty of evidence, as we have seen, that the molecules of liquids move about at high speeds, colliding with one another and with the walls of a container. If a vessel of water is left uncovered, the water gradually disappears. Some of the molecules near the surface have sufficient kinetic energies to enable them to break away from the inter-molecular attraction, plunge through the elastic film (surface tension), and escape into the air. Some of those just getting into the air may collide with gas molecules already there and be bounced back into the liquid (Fig. 11-2). If we remove some of the air above the liquid, the liquid molecules find it easier to escape as they are not bumped back into the liquid as readily. The process by which molecules escape from a liquid is called *evaporation* or *vaporization.*

If we close the vessel by means of a tightly fitting lid so that the vapor cannot escape (Fig. 11-2), some of the molecules escape from the liquid increasing the density of the vapor above the liquid. As some molecules escape from the liquid, others from the vapor return to it, until eventually these two opposing activities are proceeding at the same rate; that is, dynamic equilibrium is established. The space above the liquid is then said to be saturated. The evaporation and condensation rates are equal.

When a vapor in contact with its liquid is saturated, the rates of evaporation and of condensation are equal.

The rate of evaporation is quite variable and is affected by several conditions:

1. The temperature of both the liquid and the atmosphere.

2. The nature of the liquid. Some liquids, such as oil, evaporate very slowly; others, like ether and alcohol, evaporate quickly. Liquids which change rapidly to gases at ordinary temperatures are called *volatile liquids.*

3. The pressure upon the exposed surface of the liquid.

4. The degree of saturation of the air or space above the liquid.

5. The rate of circulation of the air over the surface of the liquid.

6. The area of the surface of the liquid exposed to evaporation.

Boiling – Usually evaporation of water occurs at the upper surface only, but at a certain temperature bubbles are produced inside the liquid, and we say it boils. Thereafter, the temperature remains constant, regardless of the rate of heating.
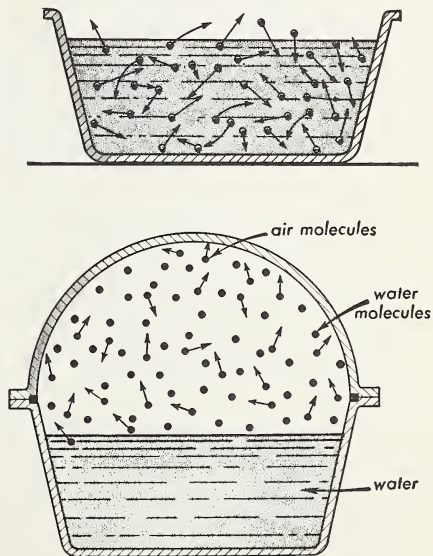

Fig. 11-2. Top: the evaporation of water from an open vessel. Bottom: the saturation of air with water in a closed vessel.

When a kettleful of water is heated, the vapor pressure of the water increases, as heat is a form of energy. *The vapor pressure of a liquid is generally defined as the pressure exerted by a vapor in equilibrium with its liquid.* The molecules are energized and so their velocities are increased. More and more of the water molecules escape from the liquid with greater and greater speeds. Finally, the vapor pressure will be equal to the atmospheric pressure, and the water begins to boil. Boiling occurs when vapor is formed throughout the liquid as well as at the top. This can take place only when the vapor pressure of the water is as great as the pressure of the atmosphere. During boiling, bubbles of vapor can be seen rising through the water. Once the boiling point has been reached, the vapor pressure stops increasing.

After the boiling point has been reached, adding more heat does not increase the temperature of the water. Additional heat is used to vaporize the water, that is, to separate the molecules from each other against the attractive forces that hold them together in the liquid.

Boiling may be defined as the process of changing from the liquid state to the vapor state, the term being restricted to the case in which the vapor pressure of the liquid is equal to the atmospheric pressure above the liquid.

OSMOSIS

Osmosis is a phenomenon observed both in liquids and gases. It is, without doubt, one of the most important natural forces. Upon it depend all forms of animal and plant life for respiration and assimilation of food. Most of us are familiar with the fact that if a dried fruit with its skin unbroken is placed in water, it will swell up almost to the original size (Fig. 11-3). This change is due to osmosis. The inside of the fruit contains a heavy sugar solution. When we placed the dried fruit in the water, the fruit gradually regained its original size because the water passed easily through the skin. After passing through, the water dissolved the sugar, thus forming a liquid inside the skin denser than the water outside. Now this
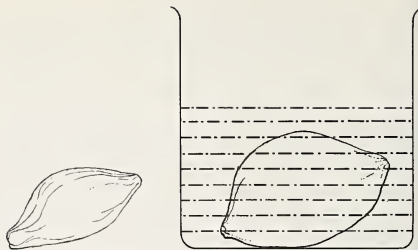

Fig. 11-3. A dried fruit swells up almost to original size when placed in water.

sugar is not able to pass out through the skin as fast as the water outside enters, perhaps because of the larger sizes of the sugar molecules. A membrane which will allow one substance to pass through it but retains another one is called a *semipermeable membrane.*

Osmosis (from Greek *osmos,* meaning "a push") *may be defined as the exchange which takes place between two fluids of different densities when separated by a moist semipermeable membrane.* The rate of exchange is always faster from the weaker solution to the stronger one.

COHESION AND ADHESION

In a liquid the molecules are much closer together than they are in gases, and do not have the freedom of motion that molecules of gases have. They do, however, have more freedom than the molecules of solids. Because they are attracted to one another with considerable force, molecules in a liquid are restricted in their motion to sliding over one another.

The attraction between like molecules is called *cohesion,* and the attraction between unlike molecules is called *adhesion.* Chalk clings together because of cohesion, but chalk clings to the blackboard by adhesion. If you dip a glass rod into a vessel of water, some of the water will cling to the glass. This shows that the attractive forces between the water molecules and those of the glass are greater than those between the adjacent water molecules. If the glass rod is dipped

into a vessel of mercury, none of the mercury molecules cling to the glass. Here the inter-attraction of the mercury molecules for one another, their cohesion, is greater than the adhesion of the mercury molecules to the glass.

SURFACE TENSION

Mercury spilled on a table top forms numerous globules of different sizes. The smaller droplets are nearly spherical and larger ones are flattened (see Fig. 11-4). This flattening is caused by the pull of gravity on the mercury. If gravity did not act, each droplet would be a perfect sphere. This

Fig. 11-4. Mercury globules are held together by cohesion.

tendency to form spherical droplets is caused by the surface of a liquid contracting and becoming as small as possible. For a drop of given volume of any liquid the surface has the least area when the drop is a sphere. This tendency of a surface to contract is called *surface tension.*

If we touch a mercury globule with a needle, a dimple is formed at the point of contact. This dimple disappears when the needle is removed. The drop behaves as though it were enveloped in an elastic membrane or surface film. Elastic films like this appear at the surfaces of all liquids. A needle or a safety razor blade can float on a water surface (Fig. 11-5). The needle or blade must be supported by the film since the density of the material is greater than that of water. The needle or blade quickly sinks to the bottom when the film is ruptured.

If you tie a loop of thread on a hoop of wire and dip the hoop into a soap solution, a film is formed over the hoop which holds the thread in place (Fig. 11-6). If the film is now punctured inside the loop, the film on the outside which remains intact draws the thread out to form a uniform circle, and behaves much as if it were a sheet of stretched rubber. The fact that the loop is uniformly curved shows that the tension of
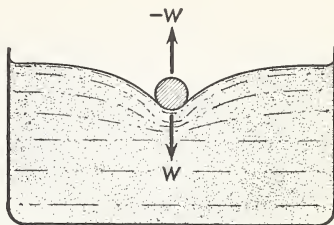

Fig. 11-5. Surface tension allows a steel needle to float by itself on water.

the film is the same in all parts of the surface. The force of surface tension is caused by the cohesive forces between molecules. The molecules inside the liquid strongly attract the molecules at the surface, and this attraction produces the effect of surface tension.

CAPILLARITY

Whenever the surface of a liquid is brought into contact with a solid, a change occurs in the level of the surface of the liquid. We are familiar with the tendency of water and other liquids to rise in minute tubes through tiny openings. Examples of this phenomenon are the rise of oil in a lamp wick, the absorption of ink by blotting paper, and the rise of sap in a tree. This behavior is called capillarity, from the Latin word *capillus,* meaning "a hair." It is explained by the relative values of cohesive and adhesive forces.
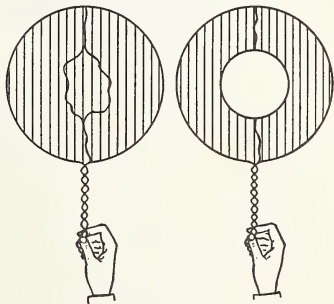

Fig. 11-6. If the soap film inside the loop is punctured, the thread forms a circle, indicating that surface tension is the same in all parts of the film.

Examine some water in a clean tumbler. You will notice that the water appears to rise along the glass surface. This is because the adhesive forces, being greater than the cohesive forces, tend to pull the water over closer to the glass walls. If mercury is poured into a glass tumbler, the surface along the glass wall dips down. This is because the cohesion of mercury molecules is greater than their adhesion to glass, and the mercury tends to draw itself together into a giant globule and so is pulled away from the glass at the upper edge of the liquid (Fig. 11-7).

Dip a glass capillary tube into a vessel of water. Note that the liquid rises in the bore because the adhesion of water to glass is greater than the cohesion of water (Fig. 11-7). Dip the tube into mercury. Note that the level of mercury inside the tube is below the outside level. The cohesive forces of the mercury molecules pull the mercury down (Fig. 11-7).

Capillarity may be defined as the rise or fall of the liquid level in capillary tubes above or below the level outside the tubes.

Capillary action appears to obey two chief laws:

1. In a capillary tube the liquid surface is higher than the outside surface if the liquid wets, that is, adheres to, the tube. The surface inside is lower if the liquid does not wet the tube.
2. The rise or fall of the surface inside the tube is inversely proportional to the radius of the tube.
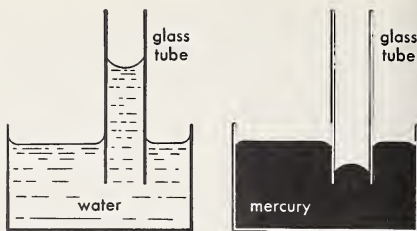

Fig. 11-7. Note that the meniscus of water is concave, that of mercury convex. Water will climb in a glass tube until the weight of the column of water balances the surface-tension force of the water in the dish. Mercury will push down until it balances the pressure of the mercury raised in the dish.

Capillary action is of great natural and practical importance. We use it every day in absorbing liquids, for example, in drying our hands and faces after washing, in wiping dishes, and in blotting ink. Lamps and candles are kept burning by the capillary rise of oil or melted wax up the spaces between the fibers of the wicks. Water is lifted from the great store of underground water to the roots of the plants by the capillary rise between the particles of the soil. A farmer cultivates a plowed field frequently if there is danger from a drought. By doing so, he increases the sizes of the spaces between the soil particles, with the result that the moisture below the surface is conserved.

THINGS TO REMEMBER

A *molecule* is the smallest particle of any substance that has the properties of the substance.

All matter is made up of molecules and the molecules are always in rapid motion.

The process by which molecules escape from a liquid is called *vaporization.*

The rate of evaporation is affected by several conditions.

Osmosis is the exchange of molecules which takes place between two fluids of different densities when separated by a moist semipermeable membrane.

Cohesion is the attraction between like molecules.

Adhesion is the attraction between unlike molecules.

Surface tension is the tendency of a surface to contract.

Capillarity is the rise or fall of the liquid level in capillary tubes above or below the level outside the tube.

QUESTIONS

1: Explain why liquids diffuse less rapidly than gases.
2: Why do clothes on a line dry sooner on a windy day?
3: What is the essential difference between vaporization and boiling?
4: The rate of evaporation depends upon several factors. Name them.
5: Define a volatile liquid.
6: Define (a) cohesion, (b) adhesion.
7: (a) Why does water wet glass?
   (b) Why does mercury not wet glass?
8: What behavior does surface tension impart to a liquid surface?
9: Why is it possible to float a needle or razor blade on water, even though the density of both the needle and the razor blade is greater than that of water? Why does greasing the needle make it easier to float?
10: A column of water in a glass tube has a concave meniscus. Why? What is the shape of the meniscus of a column of mercury? Why?
11: Why does a drop of free liquid assume a spherical shape?
12: What is capillarity? Upon what properties of a liquid does it depend? Why does it vary with temperature?
13: Why does the cultivation of the top layer of the soil help to prevent loss of moisture from the lower layers of the soil?
14: Explain how osmotic pressure, capillarity, and evaporation help in making the sap rise.
15: Can you carry water in a sieve? Try a kitchen strainer that has been well greased. Explain.
16: Why can insects walk freely on the surface of water?
17: What is osmosis? Osmotic pressure?
18: State several examples of osmosis.
19: Two toothpicks, floating in a water surface, are parallel to each other and a small distance apart. A hot needle is touched at the water surface between them and the two fly apart. Explain.
20: Why does a liquid rise higher in a narrow capillary tube than a wide one?
21: State several examples of capillarity.

LIQUIDS EXERT PRESSURE IN ALL DIRECTIONS

Because it has weight, water exerts pressure on any surface against which it rests. Water being a liquid, its particles are always trying to reach the lowest point possible. This causes its pressure to be exerted in every direction. Solids on the other hand exert pressure only in one direction, that is, vertically downward.

Suppose we have a tank filled with water or some other liquid. If we bore a hole in the bottom of the tank, the liquid will flow out, showing that liquids push *downward* upon the bottoms of their containers. If we bore a hole in one side of the tank, the liquid will flow out through this opening, showing that liquids exert pressure in a *sidewise* direction (Fig. 12-1). If we push a block of wood down into the tank, it will rise to the surface of the water again as soon as it is released. This shows that water exerts an *upward* force upon the block of wood. The upward force which liquids exert upon objects submerged in them causes some objects to lose all of their apparent weight, in which case they will float, and it causes denser objects to lose part of their apparent weight. Therefore, we conclude that at any point within the body of a liquid, pressure acts equally in all directions.

We have observed that a stack of bricks in a box pushes down on the bottom of the box. The bricks do not push out on the sides of the box. A brick has its own shape. It is rigid. But if water, or any other liquid, is put into a box, it not only presses against the bottom, but against the sides as well. Water has no shape of its own. It is not rigid. At a given depth, water pushes sidewise as much as it pushes downward and upward. We can easily measure the pressure of a liquid at any point by using a pressure gauge. A pressure gauge may be made by fastening an elastic membrane across the mouth of a thistle tube (Fig. 12-2). The stem of the thistle tube is then connected by a piece of rubber tubing to a glass U-tube partially filled with mercury. The mercury in the U-tube acts as a pressure gauge. If we press on the rubber membrane, the mercury rises in the open arm of the glass tube. The difference in level of the upper and the lower surfaces of the mercury is an approximate measure of the pressure on the membrane. If we immerse the thistle tube in a jar of liquid, the liquid pressure pushes on the rubber and we can observe that the farther down we push the
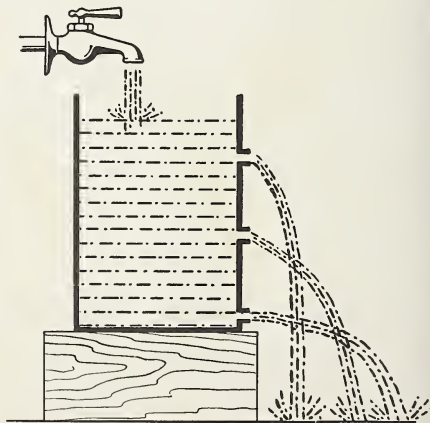

Fig. 12-1. A liquid exerts pressure in a sidewise direction.

thistle tube, the greater the pressure. Thus, we see that *the pressure exerted by a liquid is directly proportional to the depth.*

If, at a certain depth, we turn the thistle tube in different directions, we can see that the position of the mercury remains the same. This shows that *the pressure exerted by a liquid is independent of direction.*

LIQUID PRESSURE PERPENDICULAR TO THE SURFACE

At some time or other you have probably used a rather old garden hose, one that has reached the stage at which it is developing punctures, and noticed that where there is a leak the water spurts at *right-angles* to the tube.

The same point may also be illustrated by means of a small rubber ball. About half a dozen pinholes are made in it, and then it is filled with water. On squeezing it, water squirts out in all directions, all radiating from the center. This shows that *water pressure acts at right-angles to the surface.*

THE MEASUREMENT OF PRESSURE IN LIQUIDS

Just as a brick lying on a table presses upon the table, so water or any other liquid in a vessel exerts a force upon the bottom of the vessel. This force is due to the weight of the liquid. If we pile several bricks one upon another we increase the downward pressure on the surface of the table. Likewise, every layer of liquid sustains the weight of the layer directly above it. Thus we observe again the *direct relationship between pressure and depth in a liquid.* If the unit area pressed upon is 1 square foot, water 1 foot deep exerts a pressure of 62.4 pounds weight per square foot (Fig. 12-3). This is because the mass of 1 cubic foot of water is 62.4 pounds and the force of gravity upon it is 62.4 pounds weight. Similarly, the pressure at a depth of 2 feet of water is 124.8 lb. wt/sq. ft.

LIQUID PRESSURE IS PROPORTIONAL TO DENSITY

Liquids vary in density. Mercury has a density of 13.6 grams per cubic centimeter. Since the pressure caused by liquids is due to the weight of the liquids, mercury exerts a pressure 13.6 times as great as the same depth of water. Water 10 centimeters deep exerts a pressure of 10 gm. wt/sq. cm. Mercury 10 cm. deep exerts a pressure of 136 gm. wt/sq. cm. (Fig. 12-4). Since alcohol is only about 0.8 times as dense as water, the pressure exerted at a depth of 10 cm. in
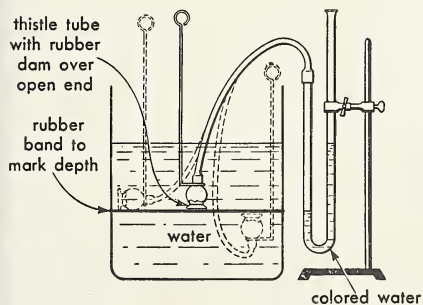

thistle tube
with rubber
dam over
open end

rubber
band to
mark depth

water

colored water

Fig. 12-2. At a given depth in a liquid the pressure is the same in all directions.
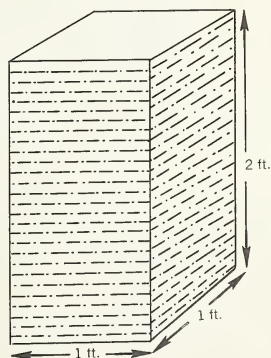

2 ft.

1 ft.

1 ft.

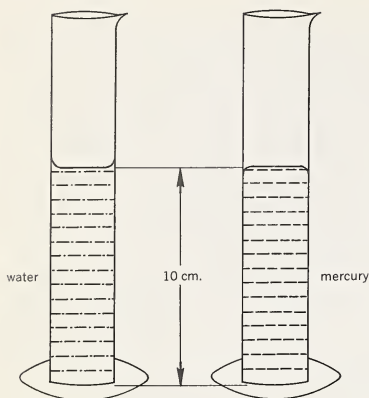**Fig. 12-3.** Water pressure at the bottom is 124.8 lb. wt/sq. ft.

Fig. 12-4. Water exerts a pressure of 10 gm. wt/sq. cm. Mercury exerts a pressure of 136 gm. wt/sq. cm.

Fig. 12-5. Liquid pressure is independent of the shape or the area of the container.

alcohol is only 8 gm. wt/sq. cm. *The pressure exerted by a liquid is directly proportional to its density.*

Liquid pressure is independent of the shape or the area of the container. Let us consider another aspect of water pressure. Suppose we have three vessels, all of the same height, the bottoms of which have equal areas, but whose shapes and volumes are not alike (Fig. 12-5). It is quite obvious that, when filled, B contains the most water and C the least. If we were to take a pressure gauge and measure the pressure at the bases of these vessels we would find that the pressures are equal. Can you explain why this is so?

Pascal used a number of different shaped glass vessels mounted in metal bases the areas of which were all the same. These were fitted into a common base. The base was fitted with a movable bottom held in position by a lever and weight. Attaching each vessel in turn to the common base and placing a suitable weight on the pan, he poured water into each vessel just to the point of forcing the bottom open, allowing some water to escape. He noted that the height of water required to open the bottom in each case was the same. Thus he proved that the liquid pressure on the bottom of a container is independent of the volume of water in the vessel and the shape of the
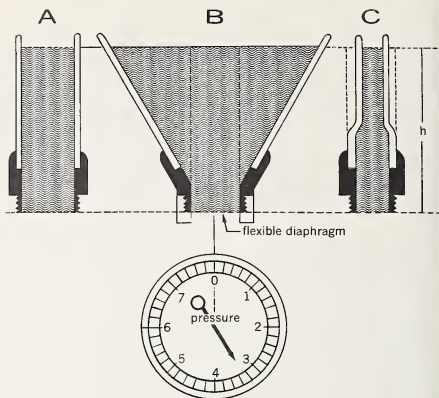
vessel, but is determined entirely by the height of the liquid above the base.

Let us take a look at a kettle filled with water. If it were true that the size or shape affect the pressure, there would be greater pressure in the main part of the kettle than in the spout, and the water would be pushed out at the spout (Fig. 12-6). We know, of course, that the levels of the water in the spout and the kettle are the same. We can, therefore, deduce that the two pressures are equal.

In conclusion, the laws of liquid pressure may be summarized as follows:

1. *Liquid pressure at a point is equal in all directions.*
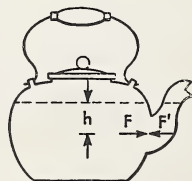2. *Liquid pressure acts at right-angles to the surface.*

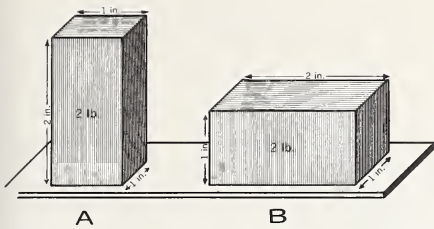Fig. 12-6. The pressures in the kettle and the spout are equal.

Fig. 12-7. A block standing on end exerts a greater pressure than when it is lying on its side.

3. *Liquid pressure is directly proportional to the depth of the liquid.*
4. *Liquid pressure is directly proportional to the density of the liquid.*
5. *Liquid pressure is independent of the area or shape of the container.*

FORCE AND PRESSURE

The word "pressure" is often used interchangeably with force; but in physics, pressure always means *force per unit area.* For example, in Fig. 12-7 the block of wood (*A*) having a mass of two pounds is resting on a table top. Since it is supported by a surface one square inch in area, the average pressure at the surface is 2 lb. wt/sq. in. The block of wood (*B*) is supported by a surface area of 2 square inches, and the average pressure is 1 lb. wt/sq. in. In each position, the force exerted against the table top is 2 lb. wt.

THE PRESSURE AT A POINT IN A LIQUID

The earth pulls down on a liquid or a gas as well as on any solid body. A vessel containing water has a force exerted from below by the water. A cylindrical glass vessel (Fig. 12-8*A*), 4 sq. cm. in cross section, is filled with water to a depth of 10 cm. Let us find the liquid pressure at the bottom of the vessel.

The bottom of the vessel supports the weight of the liquid. The volume of the liquid is 40 cubic centimeters. Since each cc. of water has a mass of 1 gram, the total mass of water is 40 grams. Hence, the downward force is 40 gm. wt. Since this downward force acts on a 4 sq. cm. area, the pressure at the bottom is 40 gm. wt. $\div$ 4 sq. cm. = $\frac{10 \text{ gm. wt.}}{\text{cm.}^2}$ This is read as 10 gm. wt/sq. cm.

Fig. 12-8 (B) represents a rectangular container 10 cm. long, 8 cm. wide, and 5 cm. high which is completely filled with water. The weight of the water is 400 gm. wt., which is the force on the bottom of the vessel. The area of the base is 80 sq. cm. The force per square centimeter, which is the pressure, is 400 gm. wt. $\div$ 80 sq. cm. = 5 gm. wt/sq. cm. The force on 1 sq. cm. of the bottom is the weight of a column of water directly over the area upon which it presses. There is a column of water 5 cm. high directly over each sq. cm. on the bottom. Each column has a volume of 5 cc.
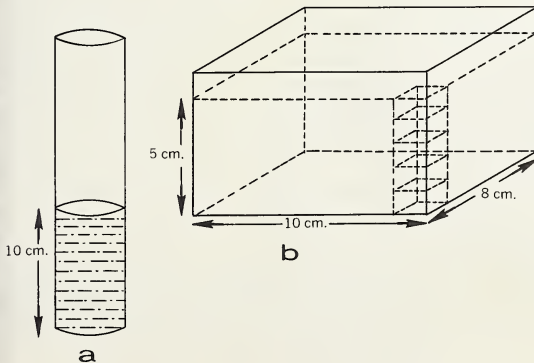

Fig. 12-8. The force exerted on each square centimeter is equal to the weight of the column of water directly over the area on which it presses.
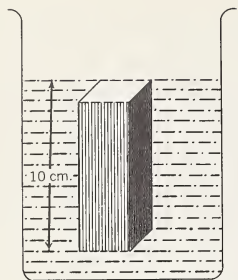
Fig. 12-9. Finding the pressure at a point in a liquid.

and the downward force is 5 gm. wt. This weight is supported on an area of 1 sq. cm., so the pressure is 5 gm. wt/sq. cm.

To find the pressure at a point in a liquid, it is helpful to imagine that a small horizontal diaphragm is supported there. In Fig. 12-9 the area of the diaphragm is 4 sq. cm., and it is 10 cm. below the surface of the water. The volume above the diaphragm is 40 cm.³, its weight is 40 gm. wt., and the pressure at the diaphragm is 10 gm. wt/ sq. cm.

The following simple rule is helpful in finding the pressure at a chosen point in a liquid.

Imagine a horizontal surface of unit area at the point. Find the weight of the column of liquid above the surface. This is numerically equal to the force per unit area, or the pressure caused by the liquid.

Example 1: What is the pressure at a point 20 cm. below the surface of water?

If a horizontal plate 1 sq. cm. in cross section were placed at this point, the weight of the water column above it would be 20 gm. wt. Therefore, the pressure is 20 gm. wt/sq. cm.

In computing the pressure produced by a column of liquid, it is helpful to use the *weight-density* of the liquid. This term means the *weight per unit volume* of a substance. For example, the weight-density of water is 1 gm. wt/cm.³ or 62.5 lb. wt/ft.³.

Thus a column of liquid of weight-density *d* and height *h* produces a pressure *p*:

$$pressure = depth \times weight\text{-}density$$
$$p = hd$$

The two quantities weight-density and mass-density must not be confused. Weight-density means weight per unit volume and is usually expressed in gm. wt/cm³ or lb. wt/ ft.³. Mass-density means the mass per unit volume and is usually expressed in gm./cm.³ or lb/ft.³.

Example 2: A diver descends to a depth of 50 m. in sea water, which has a density of 1.025 gm/ cm.³. What is the water pressure on his body?

Depth *h* is 50 m. or 5000 cm.
Weight-density *d* is 1.025 gm. wt/cm.³

Using $p = hd$

$$p = 5000 \text{ cm.}$$
$$\times 1.025 \frac{\text{gm. wt.}}{\text{cm.}^3}$$
$$= 5125 \frac{\text{gm. wt.}}{\text{cm.}^2}$$

THE TOTAL FORCE ON THE BOTTOM OF A CONTAINER

We know that the pressure which a liquid exerts on the bottom of its container equals the product of the depth and the weight-density. However, pressure is the force per unit area. In order to find the total force acting on the bottom of a container, we multiply the pressure by the entire area. In equation form

$$total\ force = area \times pressure$$
$$F = Ap, \quad \text{or}$$
$$total\ force = area \times depth \times weight\text{-}density$$
$$F = Ahd$$

Example 3: A tank 10 feet long, 5 feet wide, and 4 feet deep is full of water. What is the total force on the bottom of the tank? (Neglect air pressure.)

Area pressed upon $A = 50$ sq. ft.
Depth of water $h = 4$ ft.
Weight-density $d = 62.5$ lb. wt/ft.^{3}
   Using $F = Ahd$
                    $F = 50 \text{ ft.}^2 \times 4 \text{ ft.}$

                    $\times 62.5 \dfrac{\text{lb. wt.}}{\text{ft.}^3}$

                    $F = 12{,}500$ lb. wt.

At this time the student is well advised to use the units of measurement explicitly. That is, he should include all units in the solution and treat them as algebraic symbols as shown above. After he has become very familiar with the use of units, he may use them implicitly. If the units are used implicitly in the solution of a problem, a concluding statement is essential. The concluding statement should show the correct units. For example, the above problem may be solved as follows:

$$F = Ahd$$
$$F = 50 \times 4 \times 62.5 = 12{,}500$$

∴ the total force on the bottom of the tank is 12,500 lb. wt.

THE TOTAL FORCE AGAINST THE SIDE OF A CONTAINER

We know that a liquid exerts a force against the sides of its container. Strong tanks are often required to hold large quantities of liquids. In order to know how strong to make these tanks we must be able to calculate the force exerted on the sides. To calculate total force, we must consider the entire area pressed upon. We know that the pressure at a particular depth is the same in all directions, and we also know that the pressure

varies directly as the depth. The question is, which depth shall we choose for our calculations?

Suppose a tank 10 feet long, 5 feet wide, and 4 feet high is full of water (Fig. 12-10). We wish to find the total force exerted on one side. We know that the pressure at the surface of the water is zero and the pressure at the bottom is 250 lb. wt/sq. ft. Since the pressure varies directly as the depth, the average pressure is 125 lb. wt/sq. ft. Using the average pressure, we can find the total sidewise force by multiplying the average pressure by the total surface area. Thus, the total force on one side of the above tank is 2500 lb. wt. The formula for finding the *total sidewise force* against a vertical surface is

$$F = \frac{Ahd}{2}$$

Example 4: A swimming pool is 50 feet long and 20 feet wide. If the water is 8 feet deep, calculate the total force acting against one end of the pool.

Area of one end of pool $= 20$ ft. $\times 8$ ft. $= 160$ ft.^{2}
Depth of water $= 8$ ft.
Weight-density of water $= 62.5$ lb. wt/ft.^{3}
   Using

$$F = \tfrac{1}{2}Ahd$$
$$F = \tfrac{1}{2} \times 160 \text{ ft.}^2 \times 8 \text{ ft.} \times 62.5 \frac{\text{lb. wt.}}{\text{ft.}^3}$$
$$F = 40{,}000 \text{ lb. wt.}$$

DAMS

In building a dam, engineers must calculate the total force which the water will exert against the dam. Since the pressure varies directly as the depth, the pressure is always
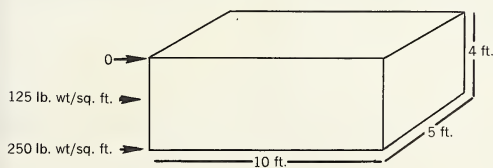

Fig. 12-10. The total force on the side of a container is the product of the area and the average pressure.
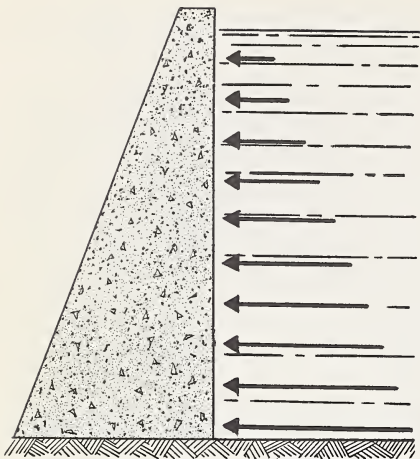
Fig. 12-11. Cross-section of a dam. Note that the increasing lengths of the arrows indicate the increasing pressure with the depth of water.

greatest at the bottom of the dam. In Fig. 12-11 notice how rapidly the thickness of the dam increases so that it can withstand the increase of water pressure with depth.

In Alberta dams are used to increase water heads for irrigation and for the operation of dynamos.

WATER SYSTEMS

The water supply for a city or a town may come from a lake in the mountains above the city level. However, in most places the water comes from a lower level, in which

case pumps must be used. The water may be pumped directly into the water mains or into a reservoir. Reservoirs are used in most of the towns and villages of Alberta. If a reservoir is used, it must be located at a height greater than the tallest building (Fig. 12-12). Since the pressure at any point in a water system is proportional to the depth of that point below the free surface, it follows that pressure depends upon elevation throughout the system. We speak of the **pressure head**, which is the vertical distance of the surface above the point in question.

PASCAL'S PRINCIPLE AND THE TRANSMISSION OF PRESSURES

Water (or any other liquid) and gases can be used to transmit pressure from one place to another. So also can a solid but in a different way. A solid can transmit pressure only in one direction, the same direction in which the original pressure is applied. Liquids and gases transmit the pressure in every direction through their volume. This phenomenon is a very important property of fluids, and one which is very useful to man.

The first man to investigate and explain the transmission of liquid pressures was Blaise Pascal (1623–1662), a French physicist. Pascal performed many experiments with vessels of different shapes. Let us consider one of them. In Fig. 12-13 we have a tank in which there are three openings, two on the top and one at the side. Each of these has the same area, and is fitted with a piston. If the box is completely filled with water and a
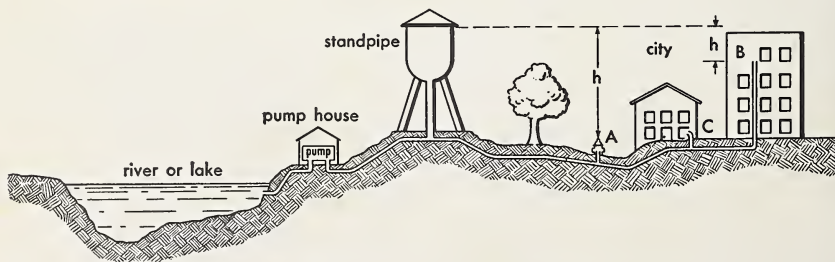

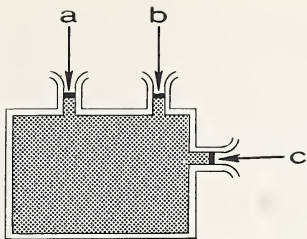Fig. 12-12. Pressure is different at various locations in a water system.

Fig. 12-13. Force applied to an area of a confined liquid is transmitted without loss to every similar area of the inside of the containing liquid.

weight placed on the piston A, it will be found that exactly the same weight must be placed upon the piston B to hold it in place. The piston C will require a force equal to the weight on A plus the force on C due to its depth. This shows that any external force applied to an area of a confined liquid is transmitted without loss to every similar area of the inside of the containing vessel.

Let us consider another vessel (Fig. 12-14). Here we have two connected cylinders having areas of cross section 1 square inch and 100 square inches respectively. Each cylinder is fitted with a piston. If a 1-pound mass is placed on the smaller piston, the piston in the other cylinder will rise.
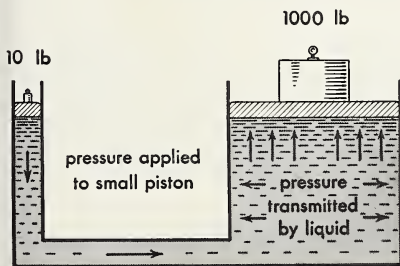
What mass must be placed on the larger piston to hold it down? Think of the last conclusion reached: any external force applied to an area of a confined liquid is transmitted without loss to every similar area throughout the body of liquid. From this we deduce that a force of 10 pounds weight on an area of 1 square inch will be transmitted without loss to each square inch of the other piston. That is, a 1000-pound mass would have to be placed upon the large piston to keep it in place. This shows that liquids may be used to multiply forces, and that the relation between the force exerted and the new force produced is in direct proportion to the areas of the surfaces.

These important facts may be summed up in what is known as *Pascal's principle:* *Pressure applied anywhere on a confined fluid is transmitted undiminished throughout the body of fluid.* Note that the principle applies to both liquids and gases.

THE HYDRAULIC PRESS

The hydraulic press is an application of Pascal's principle. Fig. 12-15 shows a simplified diagram of a hydraulic press. It consists of two cylinders, a small one and a large one. In each there is a tightly fitted

Fig. 12-14. Forces exerted on pistons vary as their areas.
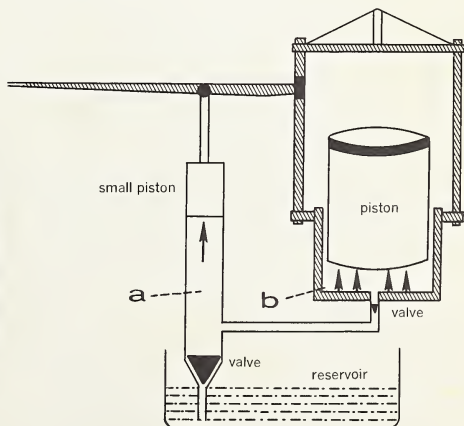

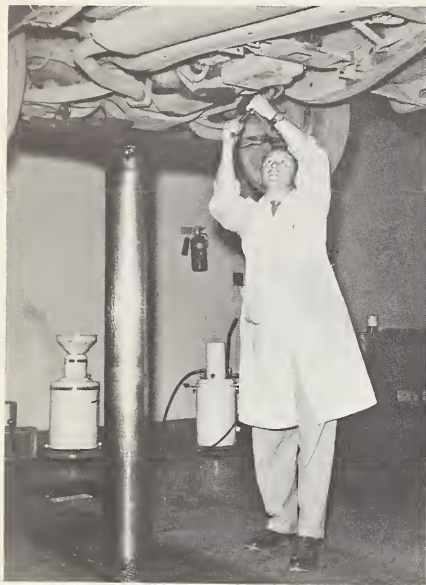Fig. 12-15. The hydraulic press is an application of Pascal's principle.

Fig. 12-16. A hydraulic car-lift.

piston. A lever is attached to the small piston. As the small piston is pushed down, some of the liquid from the small cylinder is forced into the large cylinder. This lifts the large piston through a short distance. As the liquid is pumped from the reservoir and forced into the large cylinder, it raises the large piston. The pressure may be released by opening a valve and allowing the liquid to flow back into the reservoir.

A hydraulic press is used for shaping metal parts used in machinery, lifting enormous weights, baling cotton, punching holes through steel plates, forging and die-casting, hydraulic brakes, and countless other things.

FORCE RATIO

Pascal discovered a law of pressure in liquids which, as he said, "makes it possible to multiply force to any extent whatsoever."

We often use machines to multiply the force that we can exert. Pascal's principle is often applied in machines which are used to exert great forces. In Fig. 12-14, we noted

that a force of 10 pounds weight acting on a cylinder having a cross-sectional area of 1 square inch can overcome a force of 1000 pounds weight acting on a cylinder having a cross-sectional area of 100 square inches. Since a force of 10 pounds weight is multiplied by 100, we say that the force ratio of this machine is 100 because the weight supported is 100 times as great as the force applied. Let us call the weight supported (resistance overcome), R, and the force applied, F. Then the *force ratio equals weight supported divided by force applied.*

Using F.R. for force ratio, we have

$$F.R. = \frac{R}{F}$$

Since we already know that the force exerted and the new force produced are in direct proportion to the areas of the surfaces, we can say that

$$\frac{\text{Force on large piston}}{\text{Force on small piston}} = \frac{\text{Area of large piston}}{\text{Area of small piston}}$$

If we let R represent the force on the large piston, F the force on the small piston, A the area of the large piston, and a the area of the small piston, the equation becomes:

$$\frac{R}{F} \equiv \frac{A}{a}, \quad \text{or} \quad \frac{R}{F} = \frac{D^2}{d^2}$$

since the areas are proportional to the squares of their diameters.

Example 5: In a hydraulic press the diameter of the large piston is 10 inches and the diameter of the small piston is $\frac{1}{2}$ inch. If a force of 20 pounds weight is applied to the small piston, find (a) The force on the large piston, (b) The pressure on the large piston.

1. Using $\dfrac{R}{F} = \dfrac{D^2}{d^2}$, we get

$$\frac{R}{20 \text{ lb. wt.}} = \frac{(10 \text{ in.})^2}{(\frac{1}{2} \text{ in.})^2}$$

$$R = \frac{100 \text{ in.}^2}{\frac{1}{4} \text{ in.}^2} \times 20 \text{ lb. wt.}$$

$$= 8000 \text{ lb. wt.}$$

2. Since pressure is the force per unit area, we use

$$P = \frac{R}{A}$$

$$P = \frac{8000 \text{ lb. wt.}}{22/7(5)^2 \text{ in.}^2} = 101.8 \frac{\text{lb. wt.}}{\text{in.}^2}$$

Let us look at the pressure on the small piston.

$$P = \frac{F}{a}$$

$$P = \frac{20 \text{ lb. wt.}}{22/7(\frac{1}{4})^2 \text{ in.}^2} = 101.8 \frac{\text{lb. wt.}}{\text{in.}^2}$$

From Pascal's principle we knew that the pressure applied at the small piston had to be the same as the pressure at the large one.

THINGS TO REMEMBER

Pressures in liquids at rest are:
1. exerted in all directions
2. perpendicular to the surface
3. proportional to the depth
4. proportional to the density
5. independent of the area or shape of the container.

Mass-density is the mass per unit volume.

Weight-density is the weight per unit volume.

Pressure is a force per unit area.

Pressure at a point = depth × weight-density.

Force = area × pressure.

Pascal's principle states that the pressure applied anywhere on a confined fluid is transmitted undiminished throughout the body of fluid.

Force ratio $= \dfrac{\text{weight supported}}{\text{force applied}}$

QUESTIONS

1: What is the difference between force and pressure?
2: What is meant by "head" of a body of water?
3: What specific examples show that liquids exert pressure in downward, upward, and sidewise directions?
4: The pressure exerted by a liquid depends upon two factors. Name them.
5: What unit of measurement would you use for (a) pressure, (b) total force?
6: Why does water stand at the same height in the spout as in the main part of a kettle?
7: Why is a dam so much thicker at the base than at the top?
8: State Pascal's principle.
9: List ten uses of a hydraulic press.
10: What must be sacrificed to make up for the gain in force in a hydraulic press?
11: With the aid of a simple diagram explain how the hydraulic brakes on an automobile operate.
12: Define force ratio.

PROBLEMS

1: The water level in a reservoir is 200 feet above the level of a faucet in a kitchen. Find the water pressure in the faucet if the water is not flowing.

2: Find the force acting on the bottom of an aquarium having a base 1 foot by 2 feet and containing water to a depth of 1 foot.

3: What is the pressure on the bottom of a tank of water 10 feet deep?

4: Find the total force on the bottom of a box 3 meters long, 2 meters wide, and 1 meter deep when full of water.

5: What height of water will produce a pressure of 15 pounds weight per square inch?

6: If you are swimming at a depth of 5 feet under water, find the pressure exerted by the water on your body.

7: The level of water in a standpipe is 72 feet above a faucet. Find the pressure at the faucet.

8: The water pressure on a faucet is 30 pounds weight per square inch. How high is the water level above the faucet?

9: A swimming pool is 40 feet long and 20 feet wide. If the water is 8 feet deep, find (a) the total force on the bottom, (b) the total force against one end.

10: A vertical valve in the wall of a dam is 500 feet below the water surface, and is 2 square feet in cross section. Find the force on it caused by the weight of the water.

11: Find the pressure at the bottom of a barrel 80 centimeters high if filled (a) with water, (b) with gasoline of weight-density 0.80 gm. wt/cm.³.

12: A diver descended to a depth of 290 feet in water having a weight-density of 64 lb. wt/ft³. What is the pressure at the outside of his diving suit?

13: A force of 500 pounds weight is applied to the small piston of a hydraulic press. Its diameter is 2 inches. What weight can be supported by the large piston, which has a diameter of 40 inches?

14: The piston of a hydraulic device for lifting automobiles is 6.0 inches in diameter. The device is driven by water from a city system. What is the necessary water pressure to raise a car if the total mass lifted is 3142 pounds?

15: The large piston of a hydraulic press has a diameter 20 times that of a small piston. Find the force ratio.

16: A hydraulic press has a force ratio of 100. What force must be applied to the small piston to produce a force of 3 tons weight on the large piston?

17: Find the force ratio of a hydraulic press in which the large piston has a diameter of 10 inches and the small piston has a diameter of 1 inch.

18: The diameter of a small piston of a hydraulic press is 2 inches and the diameter of the large piston is 30 inches.
   a. Find the force ratio.
   b. What effort must be exerted to produce a force of 15 tons weight?

19: The areas of the two pistons of a hydraulic press are 2 square inches and 30 square inches.
   a. Find the force ratio.
   b. What weight can be raised if a force of 15 pounds weight is applied to the small piston?
   c. How far does the large piston move if the small piston is moved 10 inches?

20: What is the diameter of the small piston of a hydraulic press if a force of 20 pounds weight on it produces a force of 4 tons weight on the large piston the diameter of which is 20 inches?

THE ATMOSPHERE

Completely enveloping the earth is a vast ocean of air which we call the atmosphere. In our daily life we are not generally conscious of this vast layer of gases. We do not notice its pressure. We move through it and breathe it with little effort under ordinary conditions. We do become aware of the pressure, however, when strong winds damage buildings and property. When we ascend or descend in an airplane we may notice the pressure change on our eardrums although many modern aircraft have pressurized cabins.

At a height of about 10 miles the atmospheric pressure is equal to 3 inches of mercury, or $\frac{1}{10}$ of the pressure on the earth's surface. At a height of 50 miles the mercury barometer would show no reading. There the vacuum is equal to that produced by a good air pump. We know that there is air 100 miles above the earth, for meteors have been seen at that height. We can see meteors only because they are heated by friction to a white heat as they plunge through the air.

The Layers of the Atmosphere – We live in the layer of the atmosphere called the *troposphere.* This extends upward from the earth's surface to a height of about seven miles. It contains about 75 per cent of the weight of the atmosphere. The temperature of the air in the troposphere decreases as we go farther away from the earth's surface, reaching about −55°F at the top of this layer (Fig. 13-1). Nearly all our clouds and storms are produced in the troposphere.

Above the troposphere is the *stratosphere.*

This region of the atmosphere extends to about 50 miles above the earth's surface. Here the air is free of almost all weather phenomena. The sky is free of clouds and dust. The sun appears very brilliant. The temperature is about −60°F.

The *tropopause* separates the stratosphere from the troposphere. It varies in height from about 55,000 feet at the equator to about 25,000 feet over the poles.

The upper part of the stratosphere is called the *chemosphere.* Most meteors burn out before they reach the chemosphere. In this region *auroras* are formed and the ultraviolet rays of the sun are filtered out. The temperature of the chemosphere is not uniform. It varies with the altitude.

Above the chemosphere is the *ionosphere* which stretches upward to about 250 miles. The temperature increases rapidly through the ionosphere. The ionosphere reflects radio waves back to the earth much as a mirror would reflect light waves, and so directly influences radio reception. Abnormal conditions in the ionosphere are capable of complete disruption of long distance radio communication.

Above the ionosphere is the *exosphere.* There is very little air in the exosphere.

Weight of Air – We are not usually conscious of the fact that air has weight. We can, however, easily weigh a given volume of air in the laboratory. For this experiment we use a Florence flask as shown in Fig. 13-2. First we weigh the flask when it is full of air. Then we pump out the air and weigh it again. We find that now it weighs less than it did the first time. The difference
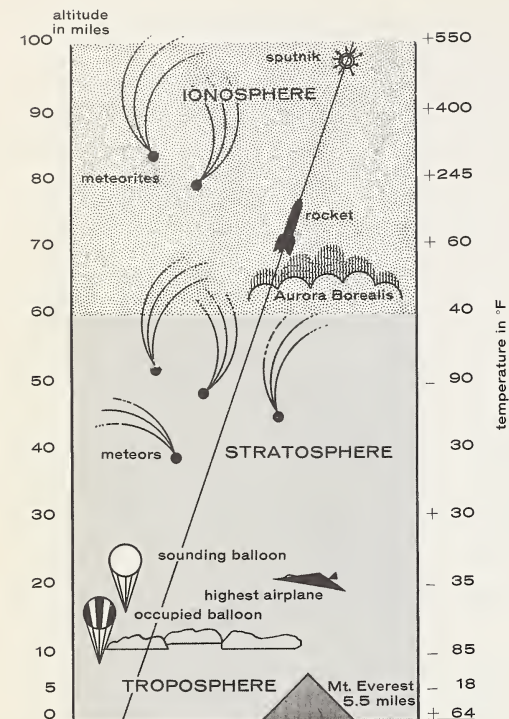
135

Fig. 13-1. An altitude chart showing the layers of the atmosphere, temperature variations, and the records set by man-made devices, probing the secret of the frontier.

in weight is the weight of the air removed. This experiment proves that air takes up space and that it has weight. *One liter of dry air at a temperature of 0°C and a pressure of 760 millimeters of mercury has a mass of 1.29 grams.*

Pressure of the Atmosphere – Since air has weight, it must exert pressure. This fact can be easily demonstrated. Stretch a rubber membrane over the upper opening of a glass bell jar open to the atmosphere. The membrane will be flat because the air pressure is equal on both sides. Put the bell jar on a pump plate, connect it to a vacuum pump, and begin to remove the air. The rubber membrane is pressed farther and farther down into the jar as the air is pumped out (Fig. 13-3). As the air from the inside of

the bell jar is removed, the upward force which the air inside the jar exerted upon the membrane is decreased, and the greater downward pressure of the air outside the jar causes the membrane to be pushed into the jar.
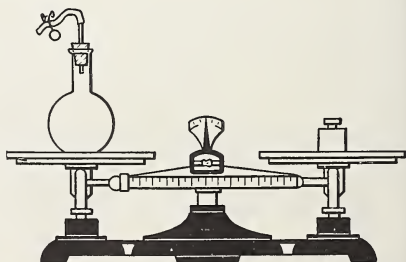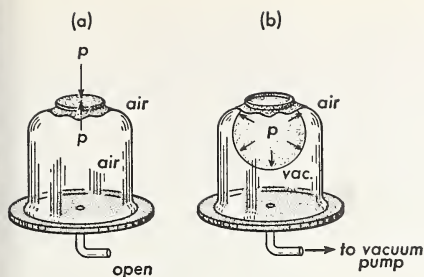

Fig. 13-2. Finding the weight of air.

Fig. 13-3. The rubber is pushed inward when the jar is evacuated.

Fig. 13-5. Magdeburg hemispheres.

If an empty varnish can is evacuated, the force of the atmosphere pushing against the sides is sufficient to crush it.

If a sheet of paper is placed over a glass tumbler brimful of water and pressed against the tumbler all around the edges, the water will remain in the tumbler even when turned upside down (Fig. 13-4). The atmospheric pressure holds the water up.

The huge forces which can be exerted by atmospheric pressures were demonstrated with the Magdeburg hemispheres by Otto von Guericke more than 300 years ago. He designed two iron vessels, each 22 inches in diameter, which could be put together to form an air-tight spherical enclosure (Fig. 13-5). When the interior was evacuated by means of an air pump, great forces were required to separate the hemispheres. Von Guericke arranged a clever public demonstration at which sixteen horses, eight on each side, pulled against each other in a tug of war, and failed to separate the hemispheres (Fig. 13-6). The force required to
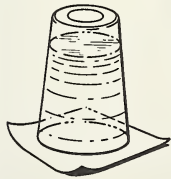

Fig. 13-4. Atmospheric pressure holds the water in the tumbler.
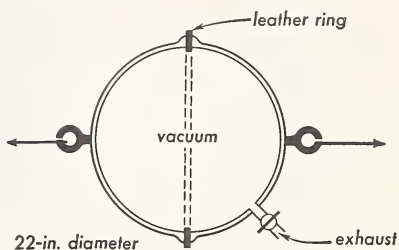
part them can easily be calculated. Replace one of the hemispheres by a flat plate. The area of the circular opening is about 380 square inches. If the atmospheric pressure is 14.7 pounds weight per square inch, the horizontal force against the plate is about 5600 pounds weight. The surface area of the hemisphere is greater than that of the plate, but the forces acting on the surface are not all in the same direction, and the total resultant force on this side is equal and opposite in direction to that on the flat plate.

WHY DO LIQUIDS RISE IN EXHAUSTED TUBES?

Whenever we sip a soft drink through a straw we show that a liquid rises in an exhausted tube. The soft drink rises in the straw because it is pushed up by the pressure of the atmosphere on the surface of the liquid outside the straw. The pressure on the surface of the liquid within the straw is reduced by suction.

More than three centuries ago, the Duke of Tuscany had a deep well dug and found that no available pump could lift the water higher than 34 feet. Evangelista Torricelli (1608–1647), a friend and follower of Galileo, was greatly interested in the reason for this and asked Galileo. Galileo replied jestingly that probably nature's abhorrence of a vacuum did not extend beyond 34 feet. This difficulty stimulated Torricelli to further endeavors, which resulted in the invention of the barometer.

Fig. 13-6. Otto von Guericke demonstrating atmospheric pressure. When air was pumped out of the hemispheres, sixteen horses, eight pulling on each hemisphere, could not pull them apart.

Deutsches Museum, Munich.

Torricelli wondered whether nature's abhorrence of a vacuum depended upon the density of the liquid used. He thought that mercury, being 13.6 times as dense as water, might be pumped to 1/13.6 of the height of water. He filled a long glass tube with mercury, closed the open end with his thumb and, inverting the tube, placed the end thus closed under the surface of a dish of mercury (Fig. 13-7). When he removed his thumb, mercury escaped from the tube until the upper surface of mercury in the tube was about 30 inches above the level of the mercury in the dish. Note that 34 feet, the height of water, is approximately 13.6 times as great as 30 inches, the height of mercury. In later experiments, it was shown that the vertical height of the column is independent of the width of the tube or its tilt (Fig. 13-8).

Torricelli's experiments proved that the height to which a liquid will rise in an exhausted tube depends upon the pressure of the air on the surface of the liquid outside the tube. The air at sea level exerts a pressure which counterbalances a mercury column 76 centimeters high, which is about 30 inches.

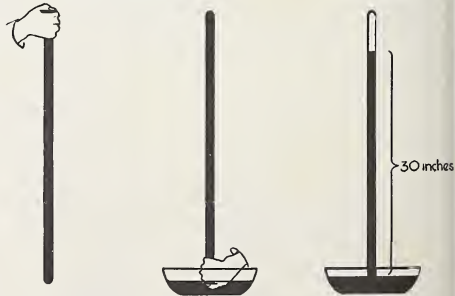***Example 1: Find the atmospheric pressure if it counterbalances a column of mercury 76 centimeters high.***

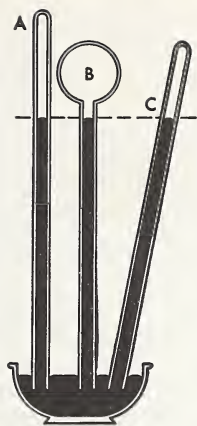**Fig. 13-7.** Torricelli's construction of the first barometer.

Fig. 13-8. The height of mercury does not depend upon the shape of the tube or its tilt.

We know that the atmospheric pressure is equal to the pressure exerted by the column of mercury.

Using $p = hd$

Mercury pressure $= 76 \text{ cm.} \times 13.6 \dfrac{\text{gm.wt.}}{\text{cm.}^3}$

$= 1034 \dfrac{\text{gm.wt.}}{\text{cm.}^2}$

In the English system the height of the mercury is 30 inches or 2.5 feet.

The density of mercury $= 62.4 \dfrac{\text{lb. wt.}}{\text{ft.}^3}$

$\times 13.6$

$= 848.6 \dfrac{\text{lb. wt.}}{\text{ft.}^3}$

$p = hd$

$p = 2.5 \text{ ft.} \times 848.6 \dfrac{\text{lb. wt.}}{\text{ft.}^3}$

$= 2121.5 \dfrac{\text{lb. wt.}}{\text{ft.}^2} \text{ or } 14.7 \dfrac{\text{lb. wt.}}{\text{in.}^2}$

THE MERCURY BAROMETER

A barometer is used to measure the pressure of the atmosphere. A mercury barometer is simply a Torricellian tube mounted in a frame and having a device for measuring the height of the mercury column. The glass
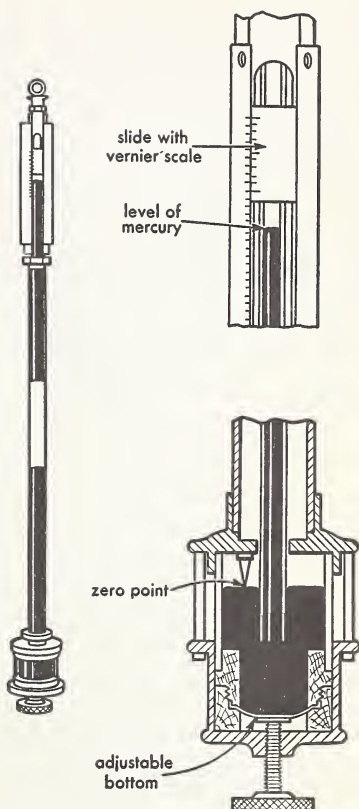

Fig. 13-9. A mechanical barometer used for accurate measurement of air pressure. The movable bottom can be raised or lowered to adjust the surface; the Vernier scale allows accurate reading.

tube and bowl are supported by a metal frame to protect them from breakage (Fig. 13-9). The scale in inches or centimeters is etched on the metal frame. The frame, tubular in form, has a vertical slot cut in it near the top, so that the level of the mercury in the glass can be seen and measured.

As the atmospheric pressure decreases, some of the mercury flows out of the tube, into the bowl. As the atmospheric pressure increases, the mercury is forced up into the tube. Since the level of mercury in the bowl varies with changes in atmospheric pressure,
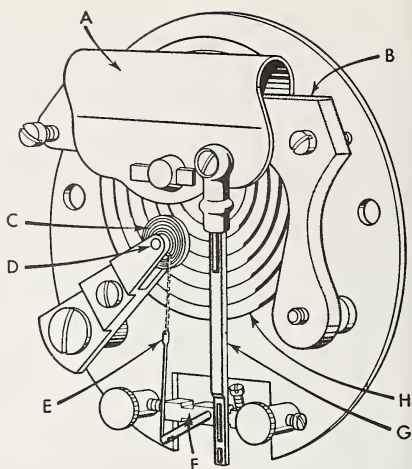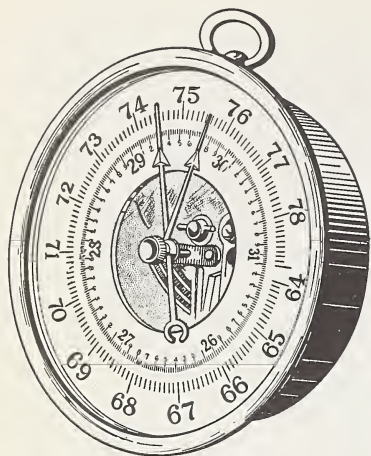
Fig. 13-10. An aneroid barometer and its mechanism.

errors in the measurements of the height of the mercury column to be read from the fixed scale will be produced. To correct for these errors, each barometer has a fixed point, usually an ivory peg, from which to measure the height of the column. This peg is mounted just inside the bowl on the frame of the barometer. The point of the peg is used as the zero mark of the scale. By means of a set screw, the surface of the mercury in the bowl can be adjusted to coincide with the tip of the ivory peg. This should be done before taking a reading.

Mercury is a suitable liquid for use in a barometer because it has a high density and a fairly low freezing point. However, mercury expands with an increase in temperature, and therefore readings taken with a mercury barometer must be corrected for changes in temperature. Mercury barometers have the advantage of great accuracy, but are not convenient to carry about. They must be fixed in a vertical position.

THE ANEROID BAROMETER

The aneroid barometer contains no liquid. It consists of a shallow box of thin, corrugated metal (Fig. 13-10). Since the air has been partially removed from the box, it is very sensitive to changes in atmospheric pressure. The base of the metal box is fastened, while the top of the box is free to move up or down in response to pressure changes. This motion, which is slight, is magnified by a system of levers connected by a fine metal chain to the spindle on which the pointer is mounted. A slight change in pressure will cause a decided change in the position of the indicating pointer. Since aneroid barometers are portable and quick in their action they are used by geologists and surveyors in measuring altitudes, in preference to the awkward, heavy, and slower-acting mercury barometers.

A self-recording barometer is known as a *barograph* (Fig. 13-11). The readings are registered on a paper scale which is moved by clockwork.

CALCULATION OF TOTAL PRESSURE

In the preceding chapter we were concerned with pressures produced by liquids, neglecting the pressure produced by the weight of the atmosphere above. For example, we found that the pressure due to the liquid at a point 100 centimeters below the surface of water was 100 gm. wt/sq. cm. This was
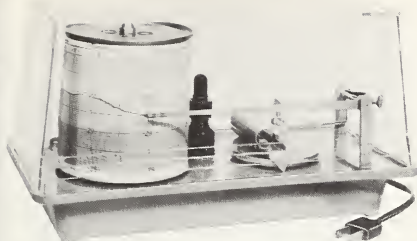
Fig. 13-11. A barograph. Pressure variations are automatically recorded over a seven-day period.

pressure in excess of the atmospheric pressure. To determine the *total pressure*, we must add the atmospheric pressure to the liquid pressure. Thus, if the barometric pressure was 75 centimeters of mercury or 1020 gm. wt/sq. cm., then the total pressure is 1120 gm. wt/sq. cm. When the excess pressure in an automobile tire is 30 lb. wt/sq. in. and the barometric pressure is 15 lb. wt/sq. in., the total pressure is 45 lb. wt/sq. in. Total pressure = liquid pressure plus atmospheric pressure.

$$P_T = P_L + P_A$$

THE OPEN-TUBE MANOMETER

The pressure of an enclosed mass of gas may be measured by an open-tube manometer (Fig. 13-12). This consists of a U-shaped tube
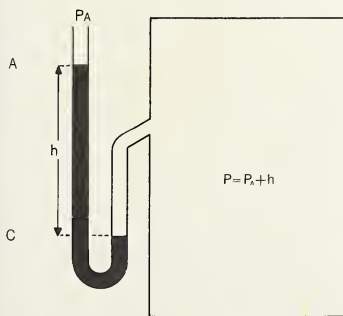
containing a liquid of known weight-density. One end of the tube is connected to the gas container, while the other is open to the atmosphere. The pressure at A in the open tube is the barometric pressure, P_A. At C there is an additional pressure, h, due to the weight of the liquid. The total pressure P is the sum of the two.

$$P = P_A + h$$

Example 2: The reading of an open-tube water manometer connected to a gas system is 10 centimeters of water, and the barometric pressure is 74 centimeters of mercury. What is the total pressure ?

$P = P_A + h$

$P = 74$ cm. of mercury $+ 10$ cm. of water

$\quad = 74 \times 13.6 \dfrac{\text{gm. wt.}}{\text{cm.}^2} + 10 \dfrac{\text{gm. wt.}}{\text{cm.}^2}$

$\quad = 1016.4 \dfrac{\text{gm. wt.}}{\text{cm.}^2}$

or 1016.4 cm. of water

THE CLOSED-TUBE MANOMETER

The open-tube manometer is inconvenient because the barometric pressure, being variable, must be determined and the correction made. The closed-tube manometer is preferable because the difference in the mercury levels indicates the correct
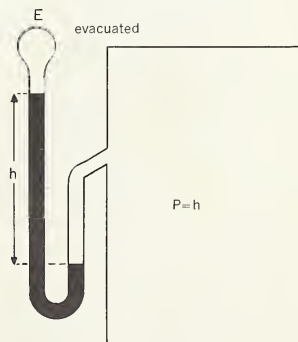

Fig. 13-12. An open-tube manometer.

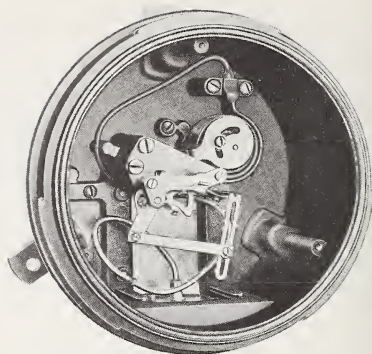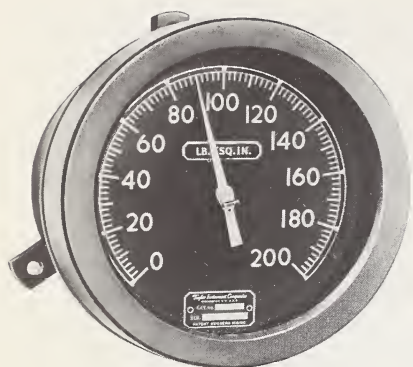**Fig. 13-13.** A closed-tube manometer.

Fig. 13-14. Bourdon pressure gauge.

pressure directly. In the closed-tube mano-meter the end, E, is sealed and the space above the column of mercury is evacuated to a high degree (Fig. 13-13). The liquid column, h, is supported entirely by the pressure of the gas in the tank. Thus the pressure of the enclosed gas is equal to that produced by this column of liquid.

THE BOURDON PRESSURE GAUGE

A pressure gauge commonly used is represented in Fig. 13-14. A flattened, thin-walled, metal tube is bent so as to lie along the arc of a circle. When the tube is evacuated, the atmospheric pressure causes the flattened walls to curl into a smaller circle. When the pressure of the enclosed air is greater than the barometric reading, the walls bulge out and the tube tends to straighten out. A gear attached to the tube rotates the pointer so that its position meas-ures the pressure. The Bourdon gauge indicates the excess of the gas pressure above that of the atmosphere. Thus, if a pressure gauge for an automobile tire reads 28 lb. wt/sq. in., the total pressure is this value plus the barometric pressure.

THE SIPHON

A siphon is simply a tube used to transfer liquids from higher to lower levels. In Fig.

13-15, two jars and a rubber tube are used. One of the jars is filled with water and the tube is placed so as to connect the two jars for the transfer of water to the empty jar at a lower level. If the tube is first filled with water and then placed in the jars, water flows from the higher to the lower level. It will continue to flow until the levels of water in the two jars are the same. If the second jar is raised to a level higher than the first, the water will flow back into the first jar until the two levels are again the same. To
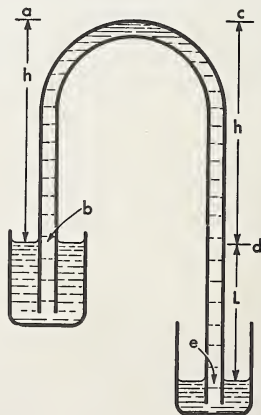

Fig. 13-15. Because there is a difference of pressure between the outlet and inlet levels, the water flows in a siphon.
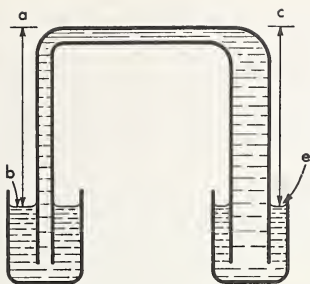
Fig. **13-16.** Water will not flow, even though the weight of water in one column is greater than in the other column.

make the siphon work, the pressure at the outlet must be lower than that at the intake.

Although the air pressure holds the water columns in both arms of the tube, the pressure of the liquid itself causes the flow. In a siphon, there are two columns of liquid, the short arm, *ab*, and the long arm, *ce*. The long arm is longer by the length, *L*. Since the downward pressure in the longer arm is greater than the downward pressure in the shorter arm, the liquid flows through the siphon. A siphon will work only so long as there is a difference between the outlet and the intake levels.

Difference of pressure makes a liquid flow through a siphon. Difference of weight has no effect. In Fig. 13-16 the column, *ce*, has a greater weight than the column, *ab*, yet the liquid will not flow because the level in the jars is the same. There is no pressure difference to make the liquid flow.

Siphons are often used for emptying vessels which cannot be conveniently tipped up, or for removing an upper layer of a liquid without disturbing the lower layer. Certain cities secure their water supply from mountain reservoirs and siphon the water over hills through large pipes called *aqueducts*.

THE LIFT PUMP

A lift pump is used to raise water from cisterns or from shallow wells. It consists of a cylinder, a piston, and two valves (Fig. 13-17). The first upstroke of the piston produces a partial vacuum in the cylinder of the pump. The push of the atmosphere forces water upward through the valve at the base of the cylinder. When the piston moves downward, this lower valve is shut by the weight of the water above the valve, preventing the downflow of the water. Meanwhile, a second valve which is in the piston itself
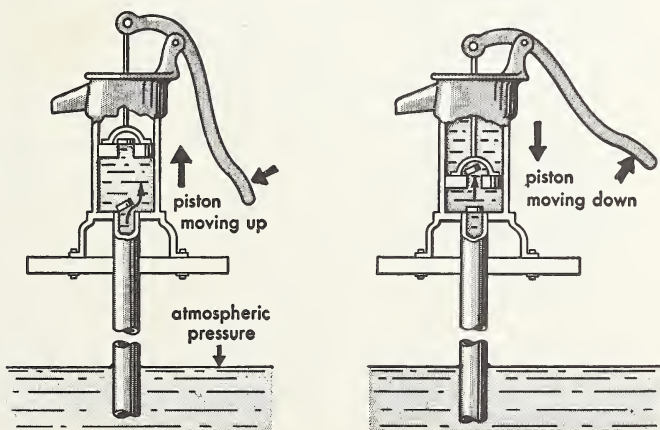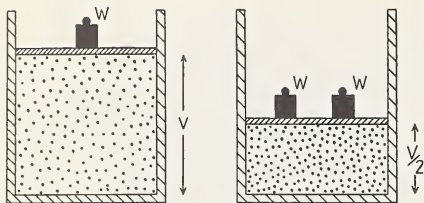

Fig. **13-17.** A lift pump.

Fig. 13-18. If the pressure of a given mass of gas is doubled, its volume is halved.

opens, and then the trapped air, followed by the water, is forced into the space above the piston. At the next upward stroke, the water which has been trapped above the piston is lifted and escapes at the spout. This type of pump will not operate if the piston is more than 34 feet above the water surface. Why?

In deep wells, the cylinder is lowered to a depth of a few feet above the surface level of the water. The water is lifted the rest of the way by the piston which is operated by a long rod. Force pumps are also used in deep wells.

COMPRESSION AND EXPANSION OF GASES

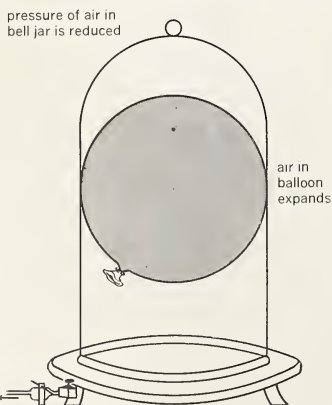Let us trap a mass of air in a cylinder by a smooth, well-fitted piston (Fig. 13-18). If

Fig. 13-19. The gas inside the balloon expands when the pressure on the outside is decreased.

we place a weight on top of the piston, the piston will move part way down. If we double the weight on top of the piston, the piston will move farther down. Under ideal conditions, doubling the pressure on a mass of gas reduces the volume of the gas to one-half.

Let us now blow up a rubber balloon, put it under a bell jar, and gradually exhaust the air that surrounds it (Fig. 13-19). As we remove air from the jar, we reduce the pressure on the outside of the balloon. The air inside the balloon expands and increases the volume of the balloon. If we let air flow into the bell jar until the pressure is the same as before, the balloon shrinks to its original size. The above experiments show that the volume occupied by a gas depends on the pressure.

The English scientist, Robert Boyle was the first to discover that there is a definite relation between the pressure and volume of a given mass of gas. In his experiments, Boyle used a J-shaped tube similar to that shown in Fig. 13-20. He poured just enough mercury into the tube to fill the bent portion. He then adjusted the mercury levels so that they would be at the same height in both arms of the tube. This mercury trapped a
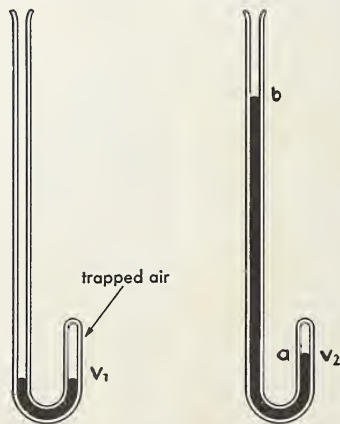

Fig. 13-20. J-tube apparatus to demonstrate Boyle's Law. At left pressure on mercury is that of air. On left pressure is that of air plus mercury column AB.

mass of air in the tube which occupied a volume, V_1. He knew that this gas must be under atmospheric pressure, because the mercury levels were the same. By reading the barometer, Boyle found the exact pressure exerted on the gas of volume V_1. Let us call this pressure P.

Next he added more mercury into the open arm of the tube until the volume of the trapped mass of air in the closed tube was reduced to one-half. Let us call the new volume V_2. He then found the new pressure, P_2, on this mass of gas by measuring the length of the mercury column, *ab*, and adding that length to the barometer reading. He found out that when the volume of the given mass of gas was reduced to one-half, the column of mercury, *ab*, was exactly equal to the atmospheric pressure. That is, the pressure on the enclosed mass of gas was exactly doubled.

As a result of several trials with this type of apparatus, Boyle found that increasing the pressure upon a given mass of gas reduced its volume correspondingly. Doubling the pressure reduced the volume to one-half. Tripling the pressure reduced the volume to one-third.

Boyle's law may be stated as follows: *The volume of a given mass of gas varies inversely as the pressure applied, if the temperature remains constant.*

Fig. 13-21 shows a convenient form of apparatus by which Boyle's law may be demonstrated. This apparatus should be placed in the laboratory of every high school. It consists of a closed glass tube and an open glass tube joined by a long flexible tube, and mounted on a stand to which is attached a meter stick. The two tubes are mounted by clamps so that they may be moved up or down for different readings. The difference in levels in the mercury and the volumes of the gas are read directly off the meter stick. Mercury is poured into the open tube, filling the flexible tube and parts of the two glass tubes. A mass of air, CD, is trapped in the closed tube. Let us now move the tubes until the level of mercury in the closed tube is the same as the level in the open tube. We know that the pressure on the enclosed mass of gas

is equal to the atmospheric pressure. Suppose the volume of the gas reads 20 cubic centimeters and the barometric pressure 30.0 inches of mercury. Since the mercury readings on the meter stick are in centimeters, we have to change the barometric pressure to centimeters.

$$\text{Thus, 30 in.} = 30 \text{ in.} \times 2.54 \frac{\text{cm.}}{\text{in.}} = 76.2 \text{ cm.}$$

Next, raise the open tube so that the mercury level in it is above the level in the closed tube and take the readings. Suppose the volume is 15 cubic centimeters and the level of mercury in the open tube is 25.4 centimeters above the level in the closed tube. Since the level of the mercury in the open tube is 25.4 centimeters higher, then the pressure on the
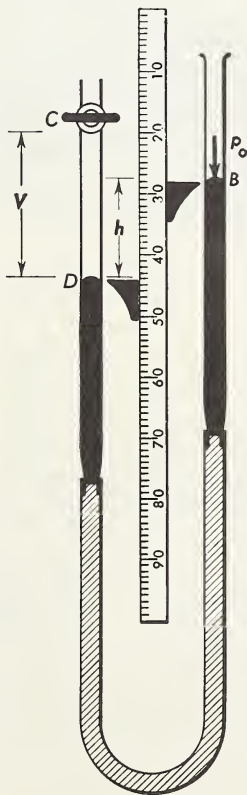
Fig. 13-21. Boyle's law apparatus.
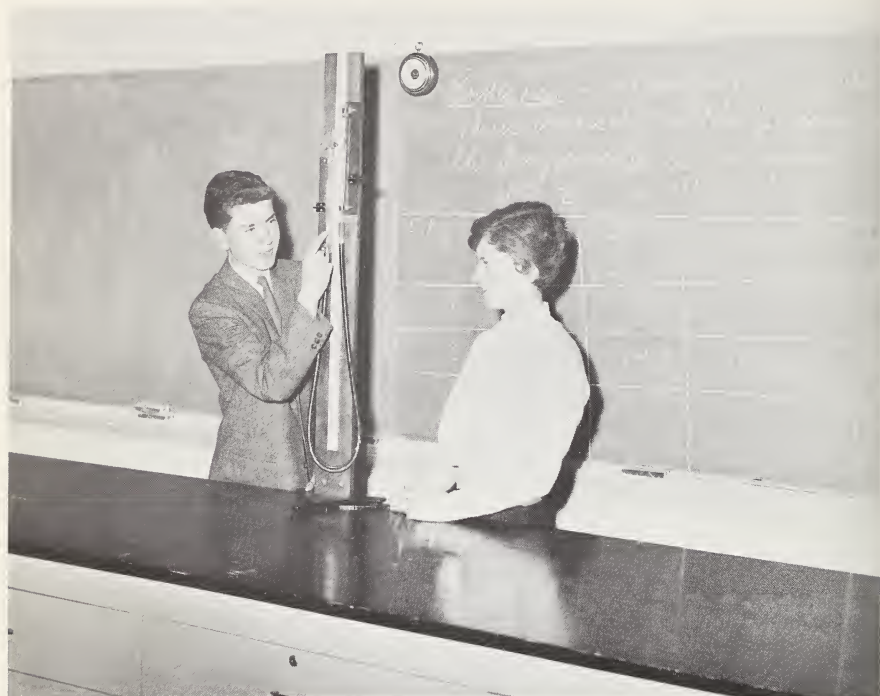
Fig 13-22. Students using Boyle's law apparatus.

enclosed mass of gas is 25.4 cm. + 76.2 cm. = 101.6 cm.

Again, lower the open tube until the level of the mercury in the open tube is below the level in the closed tube. Take the readings. Suppose the volume is 24 cubic centimeters when the mercury level in the open tube is 12.7 centimeters below the level in the closed tube. Now the pressure on the enclosed mass of air is

76.2 cm. — 12.7 cm. = 63.5 cm. of mercury.

In the accompanying chart, the results of the above experiments are tabulated. These

Experiment	Volume of air in closed arm	Level of mercury in open arm with respect to closed arm	Pressure of air in closed arm	$P \times V$
1	20.0 cc.	0 cm.	76.2 cm.	1524
2	15.0 cc.	25.4 cm. above	101.6 cm.	1524
3	24.0 cc.	12.7 cm. below	63.5 cm.	1524

show that the product of a volume of a given mass of gas and its pressure is constant. In equation form,

$$VP = k$$

or we say that

$$V_1 P_1 = V_2 P_2$$

where V_1 is the original volume, P_1 the original pressure, V_2 the new volume, and P_2 the new pressure.

Example 3: The volume of a mass of gas is found to be 500 cubic centimeters under a pressure of 750 millimeters of mercury. What volume will this gas occupy if the pressure is increased to 800 millimeters of mercury?

Using Boyle's law

$$V_1 P_1 = V_2 P_2$$

$$500 \text{ cc.} \times 750 \text{ mm.} = V_2 \times 800 \text{ mm.}$$

$$V_2 = 500 \text{ cc.} \times \frac{750 \text{ mm.}}{800 \text{ mm.}}$$

$$= 469 \text{ cc.}$$

Example 4: One liter of a certain gas has a mass of 1.50 grams at a pressure of 760 millimeters. Find its mass per liter if the pressure is decreased to 730 millimeters.

Using Boyle's law, let us find the new volume of the given mass of gas.

$$V_1 P_1 = V_2 P_2$$

$$1 \, l. \times 760 \text{ mm.} = V_2 \times 730 \text{ mm.}$$

$$V_2 = \frac{1 \, l. \times 760 \text{ mm.}}{730 \text{ mm.}} = 1.04 \, l.$$

We know that the mass of the gas does not change. Therefore, 1.04 liters of the gas at 730 millimeters has a mass of 1.50 grams.

$$\text{Mass of } 1 \, l. = \frac{1.50 \text{ gm.}}{1.04} = 1.44 \text{ gm.}$$

The mass of a substance per unit volume is known as its *density*. From the above example, it should be noted that since the pressure is reduced, the volume is increased. And, since the volume is increased and the mass is unchanged, the density is decreased. That is, the density of a given mass of gas varies directly as the pressure. The above problem may be solved as follows:

$$\text{New density} = 1.50 \frac{\text{gm.}}{l.} \times \frac{730 \text{ mm.}}{760 \text{ mm.}}$$

$$= 1.44 \frac{\text{gm.}}{l.}$$

THINGS TO REMEMBER

We live in a vast ocean of air.

Density of air is 1.29 grams per liter at 0°C and 760 mm. pressure.

Pressure of air is 14.7 pounds weight per square inch.

Liquids rise in exhausted tubes due to atmospheric pressure.

The open-tube manometer, used for measuring fluid pressures, must always be corrected for changes in the atmospheric pressure.

The closed-tube manometer indicates the correct pressure directly.

Although the air pressure holds the water columns in both arms of a siphon, the pressure of the liquid itself causes the flow.

Boyle's law states that the volume of a given mass of gas varies inversely as the pressure applied, if the temperature is kept constant.

The volume of a given mass of gas varies *inversely* as the pressure (temperature constant).

The density of a given mass of gas varies *directly* as the pressure (temperature constant).

QUESTIONS

1: Name and describe briefly the layers of the atmosphere.
2: In which layer of the atmosphere do we live?
3: Why is the ionosphere an important layer of the atmosphere?
4: What is the weight of 1 liter of air?
5: Why do we not feel the pressure of the atmosphere?
6: How can you compute the amount of pull required to separate two Magdeburg hemispheres when the system is evacuated?
7: Why does water rise in an evacuated tube?
8: Why does water not escape freely when a jug is suddenly inverted?
9: How should the reading of a mercury barometer with a large bore compare with one having a small bore? Why?
10: Why does air enter the lungs when a person inhales?
11: An automobile driver notes that the pressure gauge reads 26 lb. wt/sq. in. when he tests one of the tires of his car while at a mountain pass. Later, when he is in a valley, the tire gauge reads 25 lb. wt/sq. in. Explain.
12: What is the pressure of the atmosphere in (a) lb. wt/sq. in, (b) gm. wt/sq. cm., (c) cm. of mercury?
13: State the advantages and disadvantages of a mercury barometer.
14: State the advantages and disadvantages of an aneroid barometer.
15: What is a barograph? What is it used for?
16: What two conditions must be specified when measuring the volume of a gas?
17: State Boyle's law.
18: How is the density of a gas affected by the pressure on it?
19: What is the difference between the behavior of gases and liquids under pressure?
20: Express Boyle's law as a formula.

PROBLEMS

1: Find the volume of air in your room.
2: If 1 cubic yard of air has a mass of 2 pounds, find the mass of air in your room.
3: A flask that has a volume of 2000 cubic centimeters weighs 450.0 grams weight when the air is pumped out of it, and 452.5 grams weight when it is full of air. Find the weight of one liter of air.
4: If the air pressure is 15 lb. wt/sq. in. and the diameter of the Magdeburg hemispheres is 22 inches, find the force required to separate the hemispheres if they are evacuated.
5: If a mercury barometer reads 74 cm., what is the reading of (a) a water barometer, (b) one filled with glycerine of density 1.25 gm. wt/cc.?
6: A pair of Magdeburg hemispheres is 2.0 inches in diameter. The barometric pressure is 29 inches of mercury. What force is required to separate them if (a) the interior is completely evacuated, (b) the pressure inside is 9.7 inches of mercury?
7: What is the pressure in lb. wt/sq. in. of a pressure of three atmospheres?
8: If the barometer stands at 735 mm., what is the atmospheric pressure in gm. wt/sq. cm.?
9: If the average atmospheric pressure is 15 lb. wt/sq. in., calculate the total force on the six faces of an evacuated box 6 inches long, 4 inches wide, and 3 inches high.
10: Find the total pressure (of water and air) at a point 15 feet below the surface of a lake.
11: Calculate the combined pressure of air and water on an object immersed 30 meters under the surface of the water.
12: What is the total pressure of air and water on the body of a diver 150 feet below the surface of the water?

13: On a certain day the mercury barometer stands at 760 mm.
(a) At what height would a water barometer stand? (b) If the mass of one cubic centimeter of alcohol is 0.80 grams, at what height would an alcohol barometer stand?

14: The volume of a mass of gas is 10 cubic feet under a pressure of 2 atmospheres. Find its volume under a pressure of 1 atmosphere.

15: The volume of an inflated automobile tire is 10 liters. If the tire is to be inflated to a pressure of 2.5 atmospheres, what volume of air at one atmosphere will be required?

16: The volume of a mass of gas is 1000 cubic centimeters at a pressure of 76.0 centimeters of mercury. What pressure must be exerted upon this gas to reduce its volume to 400 cubic centimeters?

17: What pressure must be exerted on 20 cubic feet of air at 5 atmospheres in order to make its volume 100 cubic feet?

18: A gas occupies 500 milliliters at 750 mm. pressure. What volume will the gas occupy at 800 mm. pressure?

19: The barometer at the bottom of a hill reads 29.8 inches. At the top of the hill it reads 29.4 inches. Find the height of the hill.

20: The density of sulphur dioxide gas is 64 grams per liter at 760 mm. pressure and 0°C. Find its density at 800 mm. pressure and 0°C.

21: How high is an airplane when the barometric pressure is 10.0 centimeters of mercury less than at the earth's surface? The average weight-density of air is 0.00123 gm. wt/cc.

PRESSURE AND VELOCITY

In the two preceding chapters we have been discussing fluids at rest. While fluids are at rest the pressures are the same at all points at the same depth. This law no longer holds if the fluid is moving. When water flows in a uniform horizontal pipe, there is a fall of pressure along the pipe in the direction of the flow. This is because force is required to overcome friction in the fluid. In order to study frictional effects in a moving liquid, let us fill a large tank with water and connect it to a horizontal pipe fitted with tubes

(Fig. 14-1). When the valve is closed, the water seeks its level and rises to the same height in each vertical tube. When the valve is opened slightly so as to permit a small rate of flow, the water level falls in each tube, thus showing a progressive decrease of pressure along the horizontal pipe. If the valve is opened farther, so as to double the rate of flow, the pressure drop is twice as great as before. The pressure drop is proportional to the rate of flow.

Frictional effects must be overcome when fluids are transported long distances. Petroleum is forced through pipe lines from
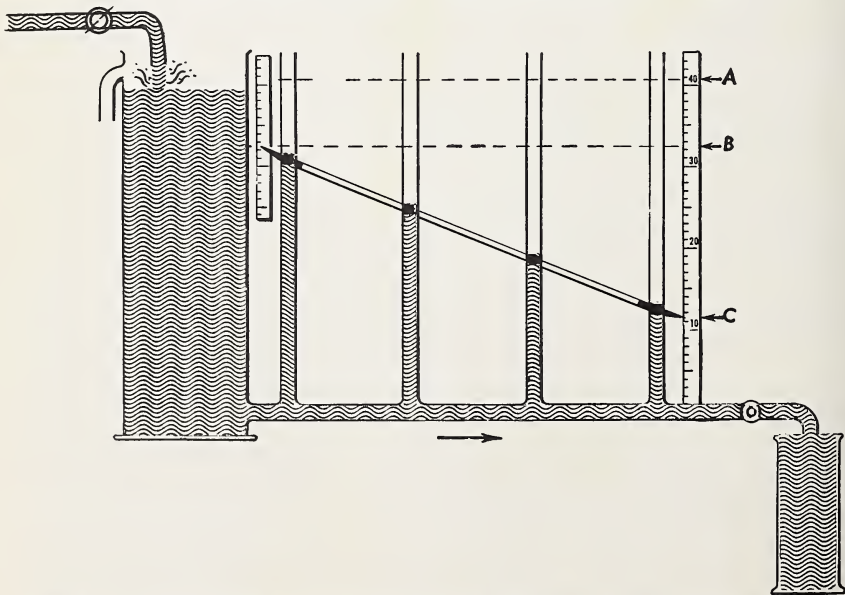

Fig. 14-1. Friction causes a fall of pressure through a tube through which a liquid flows.
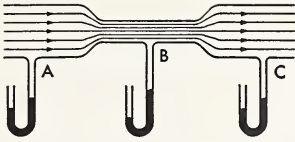
Fig. 14-2. The velocity of the fluid increases as it moves from *A* to *B* because the pressure at *A* is greater than that at *B*.

Fig. 14-3. A Venturi tube.

Alberta to Eastern Canada. Powerful pumps at the starting point raise the pressure of the liquid to several hundred pounds per square inch. If there were no flow, the pressure would be the same at all points along a horizontal pipe. Since the pressure decreases with distance because of the friction of the flowing oil, pumps are installed at a number of places along the route in order to renew the pressure and maintain the flow.

When a river flows through a wide, level plain, the water travels slowly, but in flowing through a narrow gorge its speed increases. Similarly, when water flows through a pipe which has a narrow constriction, the water speeds up as it moves through the constriction (Fig. 14-2). Hence the velocity of any molecule must increase as it moves from A to B. To cause this acceleration the pressure at A must be greater than that at B.

This illustrates a principle discovered in the eighteenth century by a Swiss scientist Bernoulli. *Bernoulli's principle* may be stated as follows:

Whenever the velocity of a horizontally moving stream of fluid increases owing to a constriction, the pressure must decrease. High velocity is associated with low pressure, and low velocity is associated with high pressure.

THE VENTURI TUBE

The Venturi tube (Fig. 14-3) provides another example of the Bernoulli principle. A fluid flows through a pipe the diameter of which differs at various points along its length. At B the cross-sectional area is much less than it is at A. If all the fluid moving past A flows through the pipe at B, then the velocity at B must be much greater than at
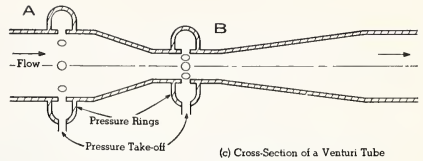
A. The velocity of the fluid at B is found by measuring its pressure. Knowing the velocity of the fluid and the cross-sectional area of the pipe, one can easily calculate the number of gallons of fluid flowing through the pipe in any interval of time.

THE INSECT SPRAYER

An insect sprayer is another application of the Bernoulli principle. The small tube extending from the liquid container is fastened to the sprayer at the exact level of the opening C. As the piston is pushed forward, a stream of air is forced across the top of the tube reducing the air pressure at C, and therefore inside the tube. The air inside the container then expands and the pressure exerted on the surface of the liquid forces it up the tube where it is broken up into a fine spray of air stream from the sprayer.

THE CARBURETOR

Another application of Bernoulli's principle is the automobile carburetor (Fig. 14-5).

Fig. 14-4. An insect sprayer. As the piston moves back and forth, the liquid rises in the narrow tube and escapes in a fine spray.

Fig. 14-5. A carburetor.

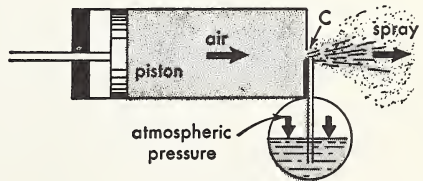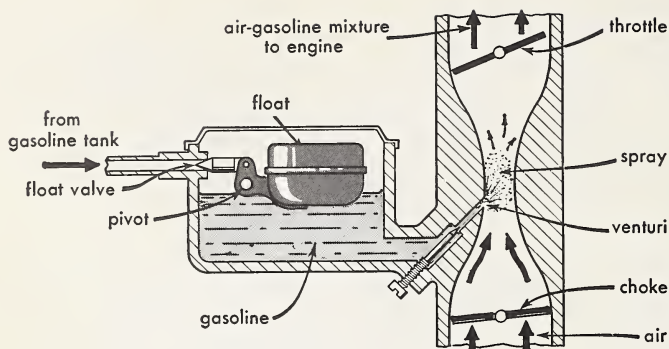The purpose of a carburetor is to mix gasoline vapor and air in the correct proportions to form an explosive mixture. Gasoline is fed from a tank to the carburetor by means of a fuel pump. The gasoline level is controlled by a small float with a needle valve attached. When the engine is running, a partial vacuum is produced in the chamber by the action of the pistons. Air sweeps past the end of the nozzle and mixes with the gasoline which is escaping. The nozzle is narrowed to give the air a greater velocity. This reduces the pressure that the air exerts. Forced through the nozzle, the gasoline vaporizes more readily in this low-pressure area. To start the engine in cold weather, a choke valve is closed a fraction, partly shutting off the air supply. In this way the pressure at the nozzle is further reduced, and more gasoline is forced into the stream, increasing the "richness" of the mixture which explodes more readily.

THE CURVING BASEBALL

A baseball which is not spinning may travel in a straight line. However, a spinning ball will move along a curved path. The ball, spinning in a clockwise direction, drags the air around with it as shown in Fig. 14-6. At A, this air current is moving with the air current set up by the forward motion of the ball. At B, it is moving against this current. The air at A moves faster, and the pressure there is reduced. Thus, the ball will be pushed along a curved line as shown.

Fig. 14-6. The path of a baseball. The pressure at A is less than at B, and the ball curves toward A.

THINGS TO REMEMBER

While fluids are at rest the pressures are the same at all points at the same depth.

When fluids flow in a uniform horizontal pipe, there is a fall of pressure along the pipe in the direction of the flow.

Bernoulli's principle states that if the velocity of a horizontally moving stream of fluid increases owing to a constriction, the pressure must decrease. High velocity is associated with low pressure, and low velocity is associated with high pressure.

QUESTIONS

1: State two laws of fluid pressure applying to fluids at rest but not to those which are moving.
2: Why does the water-flow from a faucet decrease when someone opens another faucet in the same building?
3: How does the pressure exerted by a smoothly moving fluid vary with its rate of flow?
4: Draw a simple diagram of an airplane wing. How does the shape affect the amount of lifting force?
5: Explain why a spinning baseball travels in a slightly curved path.
6: When a baseball is thrown, rotating clockwise as viewed from above, does it curve to the pitcher's right or to his left?
7: How does the velocity of a jet depend upon the depth of liquid in the tank from which the liquid flows?
8: Why does a strong wind increase the draft of a chimney?
9: Upon what factors does the efficiency of a vacuum cleaner depend?
10: Hold a thin sheet of paper horizontally in front of the lips and blow past its upper surface. What happens? Explain.
11: Although the City of Edmonton uses a force system to supply water to its residents, a huge water reservoir is located in the northeastern section of the city. What is the purpose of this water tower?

EXPERIMENT 1

15-1. To demonstrate the Brownian movement

Apparatus

Apparatus as shown in Fig. 15-1, a powerful lamp (preferably an arc lamp), a microscope.

Procedure

1. By means of the bulb, fill the chamber with smoke from the tip of a match just extinguished.
2. Place the apparatus under the microscope so as to make your observation through window.
3. Illuminate the smoke in the chamber by directing a beam of light through the window.

Observation

Conclusion

1. Why do the particles of smoke appear like so many bright point-sources of light?
2. What makes these particles dart about?

ALTERNATE EXPERIMENT

Apparatus

A glass plate, some soot or gum mastic or mercuric sulphide, distilled water, a microscope.

Procedure

1. Place on the glass plate some soot or gum mastic or mercuric sulphide in distilled water.
2. Place the plate under a powerful microscope.

3. Illuminate the glass plate with a strong light.

Observation

Conclusion

Explain your observation.

EXPERIMENT 2

15-2. To demonstrate osmotic pressure

Apparatus

A thistle tube, a sheet of moistened parchment or animal membrane, some molasses, syrup or a strong salt solution, a stand, a beaker of distilled water.

Procedure

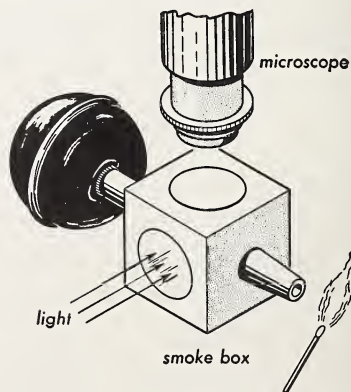1. Tie the membrane across the opening of the thistle tube.

Fig. 15-1. Experimental arrangement for demonstrating the Brownian movement.
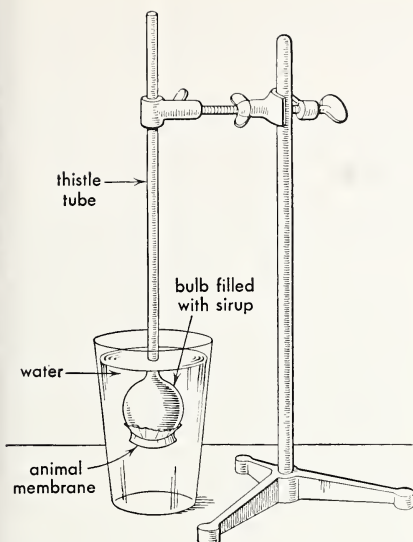
Fig. 15-2. Apparatus for demonstrating osmotic pressure.

Fig. 15-3. Experiment illustrating the forces of adhesion between glass and water.

Apparatus

A balance scale, a clean glass plate, a beaker of water, a set of weights.

Procedure

1. Attach the clean glass plate to one pan of the balance.
2. Place enough weights on the other pan to counterbalance the plate.
3. Place a beaker of water under the plate so that the plate's lower surface just touches the surface of the water.
4. Place small weights on balance pan until the plate is lifted.

Observation

1. How much force was required to lift the plate?
2. Why was the plate wet when it was pulled away?
3. Are you sure you measured cohesion and not adhesion?

Note: Water wets glass because the adhesion of water molecules to glass is greater than their cohesive force.

Conclusion

---

2. Fill the funnel and a portion of the tube with a strong solution (preferably colored).
3. Invert the thistle tube and support it in the beaker of water, so that the water outside is at the same level as the solution within the tube.
4. Let it stand for a while.

Observation

Conclusion

1. Why did the liquid rise in the tube?
2. Why did the water in the beaker become colored?
3. After about two or three hours the liquid in the tube stops rising. Why?

EXPERIMENT 3

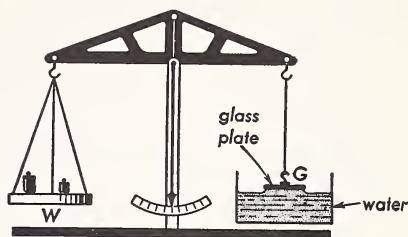**15-3. To show that force is necessary to overcome the cohesion of water molecules**

EXPERIMENT 4

To show that surface tension is the same in all directions of a film (see Fig. 11-6, p. 121).

Apparatus

A hoop of wire, a soap solution, some fine thread.

Procedure

1. Tie a loop of thread in the hoop of wire.
2. Dip the hoop into a soap solution and remove. Repeat, if necessary, until a soap film covers the entire hoop enclosure.
3. Using a needle, puncture the soap film inside the loop of thread.

Observation

Conclusion

1. What caused the loop to form a circle?
2. What does this prove?
3. What forces cause the force of surface tension?

EXPERIMENT 5

To demonstrate the change of level at the surface of a liquid when in contact with a solid (see Fig. 11-7, p. 122).

Apparatus

A set of capillary tubes or two glass plates, a shallow pan containing water.

Procedure

1. Place the capillary tubes in the shallow pan of water. Color the water with some ink.
2. If capillary tubes are not available use two plates of glass. Place the two plates of glass face to face, with a strip of wood between to keep them slightly apart along one edge while they meet along the opposite edge. Stand in shallow water.

Observation

Conclusion

1. Why did the water rise higher in the narrow part between the glasses?
2. Which is greater, the adhesive force of water molecules for glass or their cohesive force?

Repeat this experiment using mercury instead of water. Record your observations and explain your findings.

EXPERIMENT 6

To show that the pressure of water varies as the depth

Apparatus

A large glass container, a pressure gauge.

Procedure

1. Make a pressure gauge by cutting off the top of a thistle tube so that its stem is about $1\frac{1}{2}$ inches long. Connect this to a U-shaped glass tube by a rubber tube. Stretch a thin piece of rubber over the thistle tube and tie it firmly in place. Put some mercury (or colored liquid) in the U-tube and shake until its level in each branch is the same. This is a simple form of pressure gauge. Its action can be seen by pressing lightly with the fingers on the rubber membrane.
2. Carefully lower the thistle tube into the jar of water and, as it is pushed deeper into the water, notice the difference between the levels of the mercury in the branches of the U-tube.

Observation

Conclusion

1. Place the rubber membrane of the thistle tube 20 centimeters below the surface of the water.
2. Note the difference in mercury levels. Express the water pressure in centimeters of mercury.
3. Find the water pressure at this point by using $P = hd$ (in gm. wt/sq. cm.).
4. Change this water pressure to mercury pressure. (Density of mercury is 13.6 gm. wt/cc.)
5. What is the ratio of the height of water over the height of mercury? What does this prove?

EXPERIMENT 7

To show that the pressure of a liquid varies as the density

Apparatus

A pressure gauge, a tall and wide glass jar

and liquids of different densities such as water, alcohol, and carbon tetrachloride.

Procedure

1. Fill the tall glass jar with water.
2. Place the thistle tube of the gauge 2 inches under the surface of the water and record the mercury level in the U-tube.
3. Repeat at depths of 4, 8, and 12 inches.
4. Repeat experiments using alcohol and then carbon tetrachloride.
5. Tabulate your results.

Observation

Depth	Mercury reading for water	Alcohol	Carbon tetrachloride

Conclusion

1. Look up the densities of liquids used.
2. What is the relationship between the densities and the pressures at each level?

EXPERIMENT 8

To show that the pressure of a liquid at a given point is equal in all directions

Apparatus

A pressure gauge, a tall wide glass jar, water.

Procedure

1. Fill the glass jar about $\frac{3}{4}$ full of water.
2. Immerse the thistle tube into the jar of water.
3. Hold the thistle tube, membrane down, at a definite depth and record the mercury reading.
4. Hold the thistle tube, membrane up, at the *same* level as before and record the reading.
5. Hold the thistle tube, membrane side-wise, at the same level. Now turn the

thistle tube in different directions and record the readings.

Observation

Position of thistle tube		Gauge reading
1. Downward		
2. Upward		
3. Sidewise	a	
	b	
	c	

Conclusion

EXPERIMENT 9

15-4. To show that the pressure of a liquid acts at right-angles to the surface of the container

Apparatus

A glass globe and cylinder as shown, some water.

Procedure

1. Fill the glass globe and cylinder with water.
2. Exert a force on the piston and watch the jets of water thrown from all the apertures
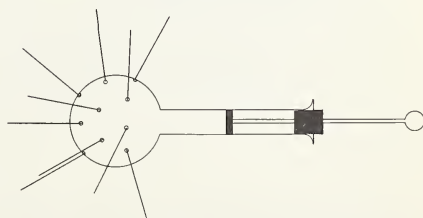
Observation

Conclusion

Fig. 15-4.

EXPERIMENT 10

To show that air has weight

Apparatus

A balance scale, a set of weights, a Florence flask, a one-hole stopper, a piece of glass tubing, a piece of rubber tubing, a pinch clamp, an air pump.

Procedure

1. Fit the flask with a rubber stopper through which a short glass tube extends.
2. Over the glass tube place a piece of rubber tubing with a pinch clamp attached.
3. Weigh the flask (full of air).
4. Remove as much air as possible and weigh the flask again.

Observation

Mass of flask filled with air = gm.
Mass of empty flask = gm.
Mass of air removed = gm.

Conclusion

EXPERIMENT 11

To show that air exerts pressure

Apparatus

An empty varnish can with a screw cap, some water, a stand, a flame.

Procedure

1. Add water to a depth of $\frac{1}{4}$ inch in the bottom of the can.
2. Boil the water in the can for a few minutes *with the top open.*
3. Remove flame and then screw the cap on tightly.
4. Let cool. Cooling may be speeded up by placing can under tap of cold water.

Observation

Conclusion

Why is it important that the can be open during boiling?

EXPERIMENT 12

To find the atmospheric pressure by using a Torricellian tube

Apparatus

A Torricellian tube, some mercury, a dish, a meter stick.

Procedure

1. Fill the tube with mercury.
2. Hold your finger against the open end of the tube and invert in a dish of mercury.
3. When the open end is under the surface of the mercury, remove finger.
4. Allow mercury to settle in tube.
5. Measure the length of the column of mercury standing *above* the mercury level in the dish.

Observation

Length of column of mercury = cm.

Conclusion

1. What supports this column of mercury?
2. The atmospheric pressure is cm. of mercury.
3. Express the atmospheric pressure in gm. wt/sq. cm.
4. Express the atmospheric pressure in lb. wt/sq. in.
5. Check by reading a barometer.

EXPERIMENT 13

To show that the volume of a given mass of gas varies inversely as the pressure applied, if the temperature is kept constant (Boyle's law)

Apparatus

Apparatus as shown on page 145, some mercury, a barometer.

Procedure

1. Adjust the open arm of the flexible U-tube of mercury so that the level of mercury is the same in both arms.
2. Measure the volume of air contained in the closed arm.
3. Read the barometer accurately and

record the atmospheric pressure in millimeters of mercury. If your barometer reads inches of mercury, change reading to millimeters.

4. Raise the open arm so that the mercury level in the open tube is above the level in the closed tube. Wait a few minutes, then take the volume of air reading. Determine the pressure to which this enclosed mass of air is subjected by *adding* to the recorded atmospheric pressure the difference in level between the mercury levels. Why do we add the difference?

5. Now lower the open arm slowly until the mercury level in the open tube is below the mercury level in the closed tube. Take the volume reading. Determine the pressure to which the enclosed mass of air is subjected by *subtracting* from the recorded atmospheric pressure the difference in level between the mercury levels. Why?

6. Repeat, taking a number of readings with the level in the open tube above that of the closed tube and a number of readings with the mercury level in the open tube below that of the closed tube.

7. Record your results.

Observation

Experiment	Volume of air	Level of mercury in open arm with respect to closed arm	Pressure of air	$P \times V$
1				
2				
3				
etc.				

Conclusion
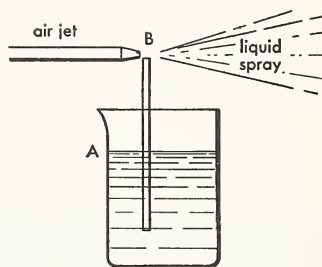
Fig. 15-5. A high velocity at B produces a lowered pressure at A.

Experiments to illustrate Bernoulli's principle

1. Connect a water tap by means of a rubber tube to apparatus shown on page 150. Open the tap allowing water to flow through apparatus. Note the water levels in the tubes. Explain.

2. Can you support a ping pong ball by a jet of air or water forced out of a nozzle? Try it with a garden hose.

3. Hold a piece of glass tubing (small bore) in a beaker of water. Blow a stream of air across the open end of the tube as shown in Fig. 15-5.

Observation
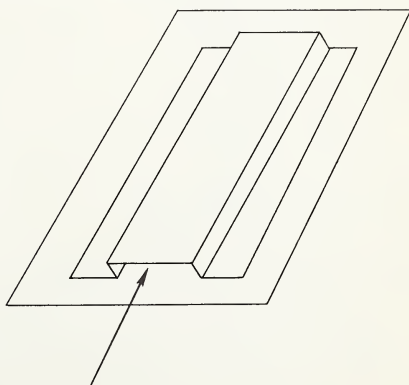
Fig. 15-6. A paper bridge 2 inches wide and 9 inches long.

Explanation

4. Cut a strip of paper about 2 inches wide and about 9 inches long. Fold to make a bridge as shown in Fig. 15-6. Place the paper bridge on a table top. Try to blow it off the table top by blowing through the bridge opening. Can you blow the paper off the table? What happens as you increase the air current by blowing harder? Explain.

5. Cut a circular piece of light cardboard about 1½ inches in diameter. Put a straight pin through its center. Place this cardboard on top of a spool. Try blowing the paper off the spool. Explain.
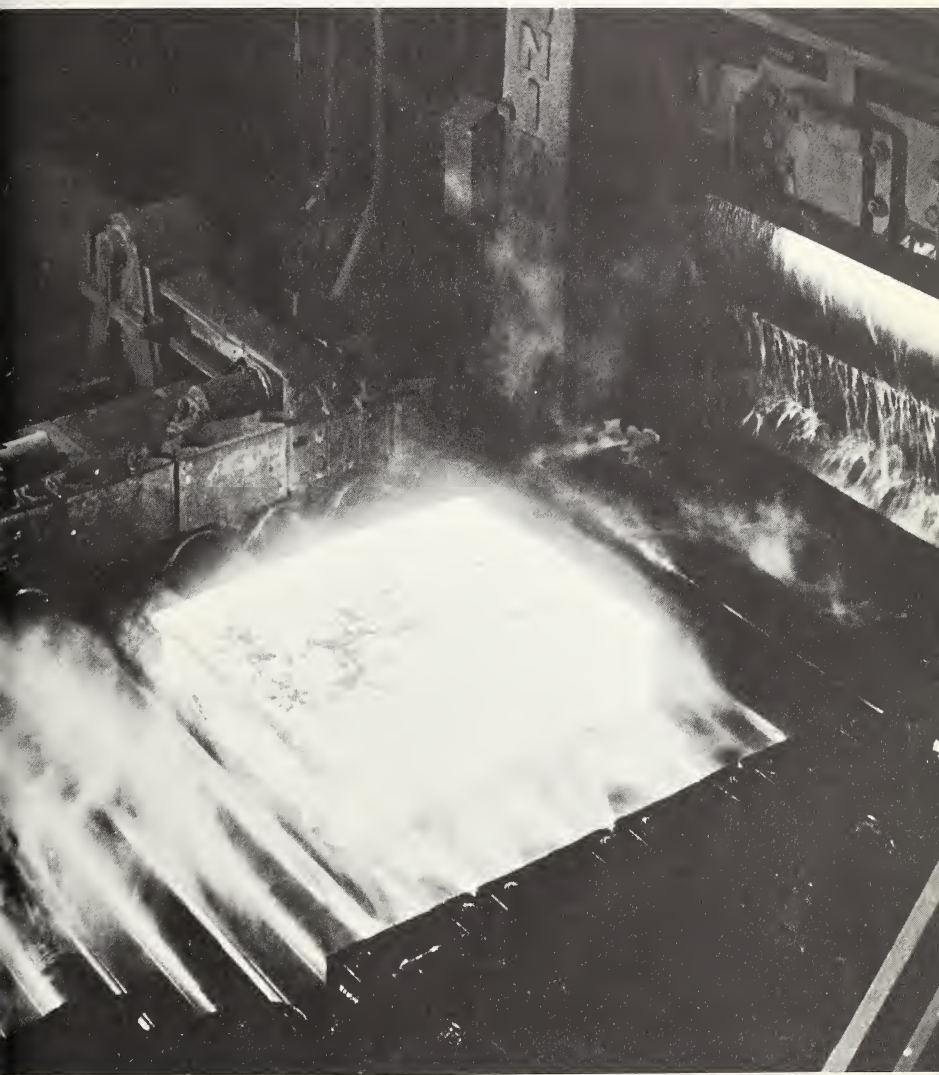
Without heat, life would be impossible. In this unit we shall study some fundamental principles of heat, and the following topics will be discussed.

Heat as a form of energy.

The difference between heat and temperature.

Units of measuring heat and temperature.

The effects of heat on solids, liquids, and gases.

Terms such as specific heat, thermal capacity, and water equivalent.

The application of heat in relation to temperature change.

A better understanding of these basic concepts will be achieved by doing the experiments outlined at the end of this unit.

NATURE OF HEAT

Many theories have been invented to explain the nature of heat. Until about the middle of the nineteenth century, heat was thought to be an invisible, weightless fluid called *caloric*. According to the caloric theory, when an object was heated in a flame large quantities of caloric were produced. It was not until the eighteenth century, when Count Rumford began to think of heat as a form of energy, that the caloric theory was abandoned. Further developments occurred in the nineteenth century, when James Clerk Maxwell (1831–1879) introduced the concept of heat as a form of molecular motion. James Prescott Joule (1818–1889), after many careful experiments, showed that the quantity of heat produced was proportional to the mechanical work done, from which it was clear that heat must be a form of energy. *Heat is the energy of molecular motion.*

EFFECTS OF HEAT ON A BODY

The most common effects produced when heat is applied to a body are (1) a rise in temperature, (2) an expansion of the body, (3) a change of state.

We know that there are exceptions to these general rules. Although most bodies experience a rise in temperature upon being heated, it is a fact that at 0°C the temperature of a block of ice remains the same until all the ice is melted. The same principle is true of boiling water. We also know that although most bodies expand upon being heated, alloys such as invar have been prepared which undergo practically no change of volume when heated, and water actually contracts if heated between 0°C and 4°C.

HEAT AND TEMPERATURE

We know from experience that heat and temperature are related. However, a difference exists between them. A cupful of boiling water and a large kettleful of boiling water have the same temperature, but they do not contain the same amount of heat. It takes much more heat to make the kettleful of water boil than it does the cupful. Finding the temperature of an object does not tell us how much heat is contained in the object. *Heat* may be thought of as the *quantity of energy* possessed by a body. This quantity of energy determines the temperature of the body.

Temperature is only relative and is difficult to define. A given substance X is at a higher temperature than substance Y, when heat tends to flow from X to Y. The word *temperature* means *intensity of heat* and may be defined as a number on a scale.

THERMOMETERS

A thermometer is a device for measuring temperatures. The first one, constructed by Galileo (1564–1642) around 1597, consisted of a bulb attached to a glass tube, the lower end of which was submerged in a vessel of water. When the bulb was heated, some of the air was expelled from it and, as the air cooled again, water was forced upward into the tube. When the temperature of the surrounding air rises, the air within the bulb

expands, forcing the water down the tube. If the bulb is cooled, the air contracts, and the atmospheric pressure forces the water up the tube. A scale attached to the narrow tube can be calibrated to any temperature scale (Fig. 16-1). Changes in barometric pressure
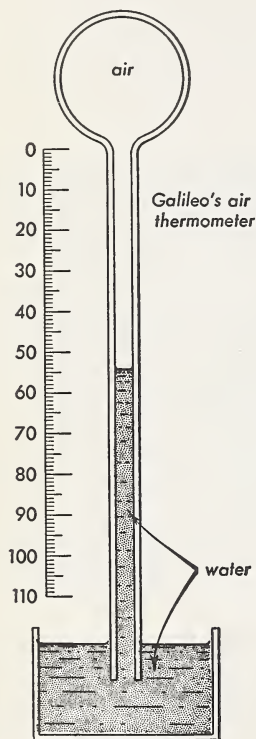

Fig. 16-1. Air on heating expands and pushes the water down in the tube. On cooling, the air contracts and the water rises.

affect the level so that a correction for such an effect must be applied to each measurement.

Of the many forms of temperature-measuring devices, the mercury thermometer is the most common. A mercury thermometer, as shown in Fig. 16-2, consists of a narrow glass tube attached to a small bulb. The bulb and part of the tube are filled with

mercury. The bulb is then heated until the mercury expands and fills the tube. The tube is then sealed off at the top. Because the air has been expelled from the tube, the mercury expands or contracts freely as the temperature changes. A scale is engraved on the glass to read temperature.

DETERMINING THE FIXED POINTS ON A THERMOMETER

The scale on a thermometer is determined by two fixed points, the freezing point and the boiling point of water. These two points have been chosen as a standard because they can be reached easily.

To determine the freezing point, the thermometer is surrounded with crushed ice (Fig. 16-3), and the lowest point to which the mercury falls is marked.

To determine the boiling point, the thermometer is surrounded by steam from boiling water (Fig. 16-4). Since the boiling point of water varies with atmospheric pressure, the apparatus must be adjusted so that water is boiled at a pressure of 760 mm. of mercury. The highest point to which the mercury rises is marked as the boiling point or upper fixed point.
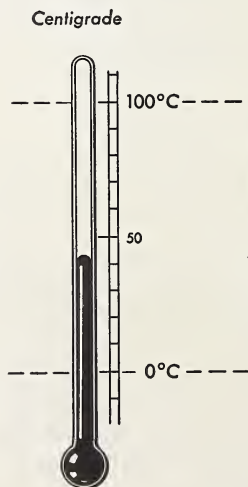
Fig. 16-2. On heating a thermometer the mercury expands more than the glass tube and the mercury level rises.
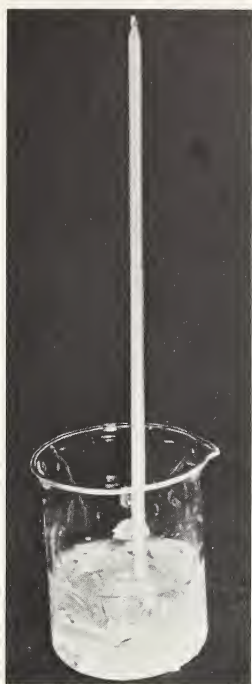
Fig. 16-3. Determination of the freezing point.

Fig. 16.4. Determination of the boiling point.
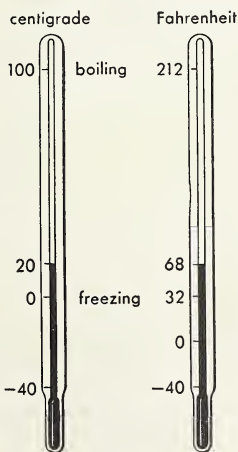
Fig. 16-5. Centigrade and Fahrenheit scales compared.

MARKING DEGREES ON A THERMOMETER

The two fixed points marked on the thermometer enable us to set up a thermometer scale. While several types of scales with degrees of different sizes have been devised, we shall study only two, the centigrade and the Fahrenheit scales (Fig. 16-5).

The centigrade scale was devised by a Swedish astronomer, Anders Celsius (1701–1744), and is often called the Celsius scale. On the centigrade thermometer, the freezing point is marked 0, and the boiling point is marked 100. The space between the fixed points is divided into 100 equal parts called degrees, and the zero of the scale is placed at the freezing point.

The Fahrenheit scale is in common use for meteorological observations and household

purposes. It was proposed by Gabriel Daniel Fahrenheit (1686–1736) of Danzig. The space between the freezing point and the boiling point is divided into 180 equal divisions, or degrees, and the zero is placed 32 divisions below the freezing point. Thus the freezing point reads 32° and the boiling point reads 212°. Fahrenheit wished to avoid negative readings and therefore chose for the zero point a temperature colder than the ordinary winter temperatures of his locality.
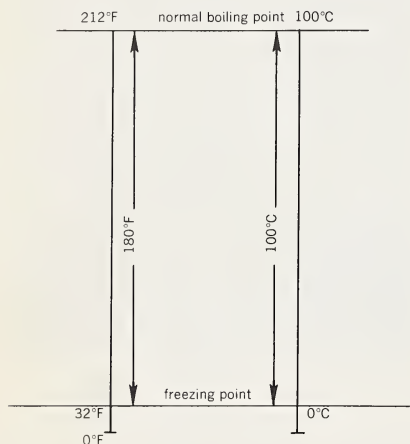
Fig. 16-6. A comparison of Fahrenheit and centigrade scales.

CONVERSION FROM ONE TEMPERATURE SCALE TO THE OTHER

In converting Fahrenheit temperature readings to centigrade, and vice versa, we must remember that 180° on the Fahrenheit scale, from the freezing point to the boiling point of water, are equal to 100° on the centigrade scale (Fig. 16-6).

$$\frac{\text{Fahrenheit degrees above freezing}}{180 \text{ F}°}$$
$$= \frac{\text{centigrade degrees above freezing}}{100 \text{ C}°}$$

or

$$\frac{\text{Fahrenheit temperature} - 32 \text{ F}°}{180 \text{ F}°}$$
$$= \frac{\text{centigrade temperature} - 0 \text{ C}°}{100 \text{ C}°}$$

or

$\frac{5}{9}$(Fahrenheit temperature − 32 F°)
= centigrade temperature

or $\quad \frac{5}{9}(\text{F} - 32) = \text{C}$

Example 1: The temperature of a room is 68°F. What is the reading on a centigrade scale?

$$\frac{68°\text{F} - 32°\text{F}}{180 \text{ F}°} = \frac{t - 0 \text{ C}°}{100 \text{ C}°}$$

$$t = \frac{36 \text{ F}°}{180 \text{ F}°} \times 100 \text{ C}°$$

$$t = 20 \text{ C}° \text{ above zero or } 20°\text{C.}$$

THINGS TO REMEMBER

Heat is the energy of molecular motion.

Heat may be thought of as a quantity of energy possessed by a body.

Temperature refers to intensity of heat.

$1 \text{ C}° = \frac{9}{5} \text{ F}°$ or $1 \text{ F}° = \frac{5}{9} \text{ C}°$. $\text{C} = \frac{5}{9} (\text{F} - 32)$ may be used to convert Fahrenheit readings to centigrade readings.

QUESTIONS

1: Define temperature difference.
2: Explain the terms *hot* and *cold*.
3: Distinguish between *temperature* and *quantity of heat*.
4: State the disadvantages of Galileo's air thermometer.
5: State the advantages and disadvantages of mercury as a thermometric substance.
6: State the advantages and disadvantages of alcohol as a thermometric substance.
7: What changes are commonly produced in a body when heated?

PROBLEMS

1: Express the following centigrade temperatures in degrees Fahrenheit: (a) 0°, (b) 20°, (c) 100°, (d) 50°, (e) 75°, (f) −40°, (g) −195°.
2: Express the following Fahrenheit temperatures in degrees centigrade: (a) 32°, (b) 68°, (c) 212°, (d) 50°, (e) 95°, (f) −40°.
3: At what temperature does a Fahrenheit thermometer read the same as a centigrade thermometer?
4: If water is cooled from 60°C to 20°C, what is the temperature change in Fahrenheit degrees?
5: To how many Fahrenheit degrees are the following centigrade degrees equivalent: 5, 15, 45, 65?
6: To how many centigrade degrees are the following Fahrenheit degrees equivalent: 9, 27, 36, 95?
7: Construct a graph for converting degrees from centigrade to Fahrenheit.

CHAPTER 17 EXPANSION OF SOLIDS

A solid body expands when heated. The heat increases the kinetic energy of the atoms, causing them to vibrate with greater average speed and push one another away. There are numerous familiar examples of expansion and contraction caused by variations in temperature. In laying a railroad, spaces are left between adjacent rails to allow for expansion. Hot water poured into a thick glass tumbler often breaks the glass, because the surface layers near the water expand more rapidly than the outer layers. There is less danger of breaking if the walls are very thin, because all parts of the glass are heated more uniformly. People sometimes place a spoon in a glass tumbler before pouring hot liquids into it. Does this serve any purpose? In constructing the steel frame for a bridge, the rivets are hammered while hot, and in cooling they draw the iron plates together.

We make use of the fact that different metals expand at different rates. If we rivet a bar of iron and a bar of brass together

(Fig. 17-1) and heat the bimetallic strip, it will bend, because the brass expands more than the iron. Many thermostats for controlling furnaces and electrical heating devices utilize such bimetallic strips.

The rate of a pendulum clock is controlled by the pendulum which swings more slowly if the length is increased. There are a number of ways to compensate for the changes in the length of a pendulum caused by temperature changes. A pendulum bob may consist of a glass cylinder filled with mercury (Fig. 17-3). If the pendulum is to swing at a constant rate, its length must remain constant. The length of the pendulum is measured from the point of suspension to

Fig. 17-1. Differential expansion.
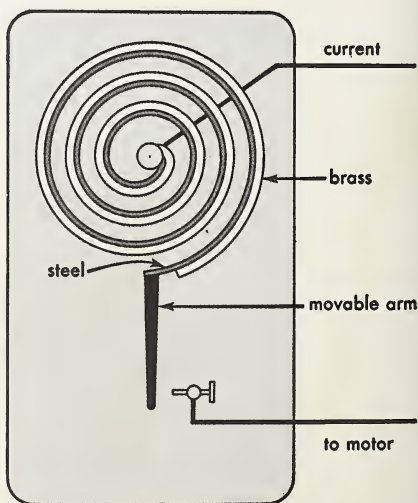
Fig. 17-2. A thermostat.

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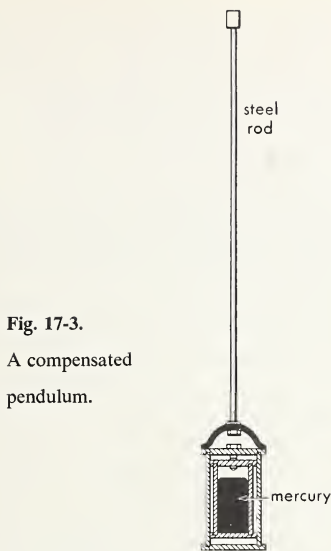
Fig. 17-3.

A compensated

pendulum.

the center of gravity. With an increase in temperature the rod expands, lowering the center of gravity. But the same increase of temperature causes the mercury to expand, thus raising the center of gravity. By properly adjusting the quantity of mercury the two opposing tendencies may be made to cancel out.

A compensated balance wheel has its rim made of two metals, brass and steel, the brass being on the outside (Fig. 17-4). With a rise in temperature the two metals expand, and this throws the weight of the rim farther out from center. Because the rim is a bimetallic strip, the expansion causes an increased amount of bending, and this brings the rim nearer to the center. By correctly adjusting the weight of the rim, the two effects may be made to cancel out. The weight of the rim may be adjusted by attaching screws to it.

COEFFICIENT OF LINEAR EXPANSION

When a metal rod is heated, the expansion is proportional to the original length (l_1) and the rise of temperature ($t_2 - t_1$). If a rod is

1 foot long and the rise in temperature is 1 centigrade degree, the increase in length is known as the coefficient of linear expansion. By using precise measuring instruments we can find the increase in the length of any rod. If an aluminum rod 1 foot long is heated one centigrade degree, its increase in length is 0.000023 feet. One foot of iron heated one centigrade degree expands 0.000012 feet. The amount of expansion for different materials of the same length will vary at the same change in temperature.

The coefficient of linear expansion may be defined as the increase in length experienced by a rod of unit length when its temperature is raised 1 degree.

The coefficient of linear expansion of aluminum is 0.000023/C°. Since the coefficient of linear expansion is defined as the increase in unit length per degree rise of temperature, its value does not depend upon the length unit used but upon the size of the degree used to measure the temperature change. The accompanying table shows the coefficient of linear expansion for centigrade degrees. The coefficient of expansion for Fahrenheit degrees is 5/9 as much.
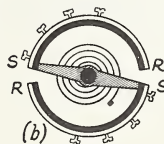

Fig. 17-4.

A compensated

balance wheel.

Coefficient of Linear Expansion

Material	Fractional expansion per centigrade degree
Aluminum	0.000023
Brass	0.000019
Copper	0.000017
Glass (ordinary)	0.0000090
Glass (pyrex)	0.0000040
Iron	0.000012
Steel	0.000013
Platinum	0.0000090
Invar (nickel-steel alloy)	0.0000009

Example 1: How much will a 10-foot length of iron expand if its temperature is raised 10 C°?

1 foot for a rise of 1 C° expands 0.000012 feet.
10 foot for a rise of 1 C° expands 0.00012 feet.
10 foot for a rise of 10 C° expands 0.0012 feet.

We conclude that the increase in the length of a solid when heated is equal to the product of its length, the temperature change, and its coefficient of linear expansion.

This may be expressed by the formula:

$$\Delta l = l(t_2 - t_1)\alpha$$

where Δl is the increase in length
l is the original length
t_2 is the final temperature
t_1 is the original temperature and
α is the coefficient of linear expansion.

Example 2: A steel rod is 50 centimeters long at 0°C. How much will it expand when heated to 60°C?

Using $\Delta l = l(t_2 - t_1)\alpha$
$\Delta l = 50$ cm.$(60 - 0)\cancel{C°} \times 0.000013/\cancel{C°}$
$\Delta l = 0.039$ cm.

Example 3: An iron rod is 60 centimeters long at 0°C, and the increase in length is 0.0528 centimeters at 80°C. Find the coefficient of linear expansion of iron.

Increase in length $(\Delta l) = 0.0528$ cm.

Using $\Delta l = l(t_2 - t_1)\alpha$

0.0528 cm. $= 60$ cm. $\times 80$ C° $\times \alpha$

$$\alpha = \frac{0.0528 \ \cancel{cm.}}{60 \ \cancel{cm.} \times 80 \ C°}$$

$$\alpha = 0.000011/C°.$$

Example 4: The main span of a steel bridge is 1800 feet long. What change in its length would occur between −25°F and 100°F?

Coefficient of linear expansion of steel is
0.000013/C°

or
0.000013 $\times \frac{5}{9}$/F° $= 0.0000072$/F°

$$\Delta l = l(t_2 - t_1)\alpha$$

$$\Delta l = 1800 \text{ ft. } [100 - (-25)]\text{F°} \times 0.0000072/\text{F°}$$

$$\Delta l = 1800 \text{ ft. } \times 125 \ \cancel{F°} \times 0.0000072/\cancel{F°}$$

$$\Delta l = 1.62 \text{ ft.}$$

We may find the coefficient of linear expansion of a metal by using the apparatus in Fig. 17-5.

A horizontal brass jacket tube is fitted at the ends with corks through which passes the metal rod, the coefficient of which we wish to find. Two small side tubes are attached to the brass jacket tube, and steam is passed through the upper and escapes
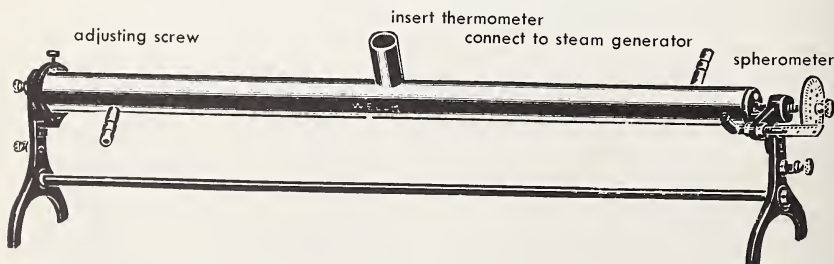

Fig. 17-5. Linear expansion apparatus with spherometer screw.

through the lower. A thermometer is fitted into a third tube at the middle.

The jacket is supported by two uprights which are fastened to the base of the apparatus. An adjusting screw touches one end of the rod, and the other end makes contact with a spherometer. The spherometer is adjusted so that the central screw just touches the end of the rod. The reading is then taken, and the spherometer screw is loosened. The temperature is recorded, after which steam is passed through the tube for a few minutes. While it is still passing through, another spherometer reading is taken. To make sure that expansion is complete, steam is passed for another two minutes. There should, of course, be no further change in the reading. The result is then worked out:

Length of rod . . . . 60 cm.
Room temperature . . . . 20°C
First spherometer reading . . . . . 2.54 mm.
Second spherometer reading . . . . 3.46 mm.
Final temperature . . . . 100°C
       Expansion = 3.46 mm. − 2.54 mm.
               = 0.92 mm. or 0.092 cm.
60 cm. of rod heated through 80 C° expanded 0.092 cm.
1 cm. of rod heated through 1 C° expanded 0.000019 cm.
∴ the coefficient of expansion is 0.000019/C°.

COEFFICIENTS OF AREA AND VOLUME EXPANSION

It must be remembered that when solids are heated they expand in all directions. Since the surface area is two dimensional, *the coefficient of area expansion, or the increase in unit area per degree, is approximately twice the coefficient of linear expansion.* Since volume is three dimensional, *the coefficient of volume expansion, or the increase in unit volume per degree, is approximately three times the coefficient of linear expansion.*

Example 5: The temperature of 100 cc. of copper is raised from 10°C to 100°C. Find the increase in volume.

Coefficient of linear expansion of copper is 0.000017/C°

Coefficient of volume expansion of copper is $3 \times 0.000017/C° = 0.000051/C°$

$$\Delta V = V(t_2 - t_1)\alpha$$
$$\Delta V = 100 \text{ cc.}(100 - 10) \text{ C}°$$
$$\times 0.000051/C°$$
$$\therefore \Delta V = 0.459 \text{ cc.}$$

∴ the volume expansion is 0.459 cc.

THINGS TO REMEMBER

Solids expand when heated and contract when cooled.

Different metals expand and contract at varying rates.

Coefficient of linear expansion is the increase in length experienced by a rod of unit length when its temperature is raised 1 degree.

Increase in length of a solid may be found by the following formula:

$$\Delta l = l(t_2 - t_1)\alpha$$

Coefficient of area expansion is approximately twice the coefficient of linear expansion.

Coefficient of volume expansion is approximately three times the coefficient of linear expansion.

QUESTIONS

1: Why do solids expand when heated?
2: What is the coefficient of linear expansion?
3: What is meant by (a) coefficient of area expansion? (b) coefficient of volume expansion?
4: How does pouring hot water on a metallic jar ring help loosen the ring?
5: Which is less likely to crack when hot water is poured into it: (a) a tumbler made of thin glass or (b) a tumbler made of thick glass? Explain.
6: When a steel plate with a hole in it is heated, does the hole become larger, smaller, or does it remain the same. Explain.
7: (a) What is a bimetallic strip?
    (b) Name three devices in which the bimetallic strip is used.
8: Why does a pendulum clock often lose time during hot weather?
9: What kind of wire would be suitable for sealing into ordinary glass?
10: Draw a diagram of a thermostat and explain how it works.

PROBLEMS

1: The coefficient of linear expansion for brass is 0.000019 for 1 centigrade degree. What is it for 1 Fahrenheit degree?
2: The coefficient of linear expansion for copper is 0.000017/C°. Find its (a) coefficient of area expansion, (b) coefficient of volume expansion.
3: A steel wire is 10 feet long at a temperature of 20°C. What is the increase in length at 25°C?
4: A brass straight edge is exactly 1 meter long at 0°C. What is the increase in length at 20°C?
5: A rod of silver is 100 centimeters long at 0°C. When heated to 100°C, the change in length is 0.189 centimeters. What is its coefficient of linear expansion?
6: A steel tape is exactly 100 centimeters long at 0°C. At what temperature will the increase in length be 0.022 centimeters?
7: A steel cube is exactly 10 centimeters along each edge at 0°C. Find the change in volume at 100°C.
8: A steel wagon tire is 14 feet in circumference at 200°C when it is put on the wagon wheel. How much will the circumference shrink in cooling to 20°C?
9: If steel rails 10 meters long are laid just touching each other, at 40°C, what will be their separation at −10°C?
10: The spaces between 40-foot steel rails are 0.20 inches at a temperature of 0°F. If the rails close up at 90°F, what is the coefficient of linear expansion?
11: A hole, 1 inch in diameter, is drilled through a steel plate at 20°C. Find the increase in diameter of the hole if the plate is heated to 100°C.
12: A steel tape is exactly 100 feet long at 0°C. A distance of 50 feet was measured off with this tape at a temperature of 20°C. What was the error?

CHAPTER 18 EXPANSION OF LIQUIDS

Like solids, liquids expand when heated and contract when cooled. They do, however, expand more rapidly when heated than do solids. This is because the force of attraction between the molecules in a liquid is much less than that in a solid.

Since liquids do not have a definite shape but take the shape of their container, we are concerned only with their expansion in volume.

The volume expansion of a liquid may be found by using a suitable apparatus as shown in Fig. 18-1. It is called a dilatometer. It consists of a bulb with a measured volume (usually 10 cc.) and a graduated capillary tube. The bulb is filled with the liquid to be tested and is heated in a water bath. Why? At various temperatures the expansion of the liquid is observed and read directly off the graduated scale. It must be remembered, however, that the container also expands when heated; therefore, the expansion read on the dilatometer scale is not the real expansion but the apparent expansion. Because of the expansion of the container the real or absolute expansion or contraction is somewhat greater than the observed or apparent expansion.

Example 1: A pyrex dilatometer containing alcohol shows a reading of 10.014 cubic centimeters at 20°C and 10.375 cubic centimeters at 50°C. Find the coefficient of absolute expansion of alcohol between these temperatures.

Neglecting the expansion of glass, 10.014 cubic centimeters of alcohol heated through 30 centigrade degrees expanded 0.361 cubic centimeters. One cubic centimeter of alcohol heated through 1 centigrade degree expanded 0.00120 cubic centimeters. Therefore, the coefficient of apparent expansion of alcohol = 0.00120/C°.

Fig. 18-1. Dilatometer—

a volume coefficient tube.

To this should be added the coefficient of volume expansion of the glass. Since the coefficient of linear expansion of glass is 0.000004/C°, the coefficient of volume expansion is 0.000004/C° × 3 = 0.000012/C°.

Therefore, the coefficient of absolute expansion of alcohol

$$= 0.00120/C° + 0.000012/C°$$
$$= 0.001212/C°$$

From the above we note that:
The coefficient of absolute expansion of a liquid is equal to the coefficient of apparent expansion of the liquid plus the coefficient of volume expansion of the container or

$$C_r = C_a + C_c$$

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Example 2: A brass gasoline tank has a volume of 15 gallons and is filled with gasoline. The average coefficient of real expansion is 0.00096/C°. What volume will overflow if the temperature rises 20 centigrade degrees?

Volume expansion of gasoline = 15 gal. × 0.00096/C° × 20 C° = 0.29 gal.
Coefficient of linear expansion of brass = 0.000019/C°
Coefficient of cubical expansion of brass = 0.000057/C°
Volume expansion of brass container = 15 gal. × 0.000057/C° × 20 C° = 0.017 gal. or 0.02 gal.
∴ the overflow = 0.29 gal. − 0.02 gal. = 0.27 gal.

Alternate solution

Since we know that the overflow is the apparent expansion of the gasoline, we can find it by using its coefficient of apparent expansion.

$$C_a = C_r - C_e$$
$$C_a = 0.00096/C° - 0.000057/C°$$
$$= 0.00090/C°$$
$$\Delta V_a = V(t_2 - t_1)C_a$$
$$\Delta V_a = 15 \text{ gal.} \times 20 \text{ C}° \times 0.00090/C°$$
$$= 0.27 \text{ gal.}$$

∴ The apparent volume expansion or the overflow is 0.27 gal.

PECULIAR EXPANSION OF WATER: ITS MAXIMUM DENSITY

Suppose we fill a dilatometer with water at room temperature. As we cool the bulb and water, the water gradually contracts until a temperature of 4°C is reached. As we lower the temperature of the water below 4°C, the water expands until its freezing point, 0°C, is reached. Conversely, it contracts when warmed from 0°C to 4°C. Above 4°C water acts like other liquids, expanding when it is heated.

Fig. 18-2 shows the peculiar behavior of water as it is cooled from its boiling point to its freezing point. This is also represented by the graph shown in Fig. 18-3. Note that at 4°C the volume has reached a minimum, and hence the density is a maximum. Thus, we say that the maximum density (mass per unit volume) of water is reached at 4°C.
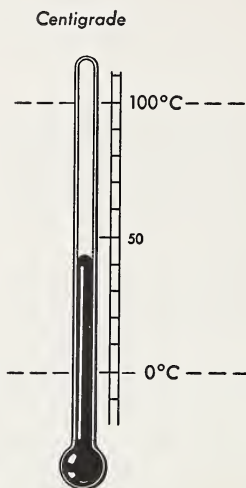

Centigrade

Fig. 18-2.

Water

expands

on cooling

from 4°C

to 0°C.

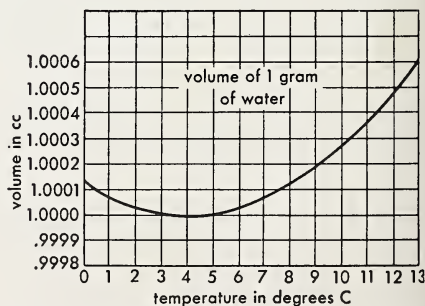In devising the metric system the gram was defined as the mass of one cubic centimeter of water at 4.0°C.

Fig. 18-3. The maximum density of water is reached at 4°C.

The fact that water is denser at 4°C than at the freezing point plays an important role

in the preservation of our aquatic life. When a lake freezes, ice is formed at its surface. This is because the surface water of a lake cools first, becomes denser and sinks to the bottom pushing the lighter, warmer water to the surface where it is cooled. This circulation continues until the whole body of water reaches a temperature of 4°C, at which point the circulation stops. Further cooling results in the lowering of temperature of the water near the surface. While the water at the bottom of a lake remains at 4°C, the water at the surface is cooled to 0°C at which point it freezes (Fig. 18-4).
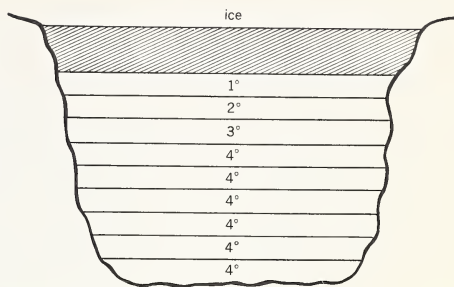
Fig. 18-4. Ice is formed at the surface because the maximum density of water is at 4°C.

THINGS TO REMEMBER

Liquids expand more rapidly when heated than solids, and to a greater extent, for a given increase in temperature.

Apparent expansion is the volume expansion of a liquid, neglecting the expansion of the container.

Real or *absolute expansion* is the true volume expansion of a liquid.

$$C_r = C_a + C_e$$

The maximum density of water is reached at 4° C.

QUESTIONS

1: Why do liquids expand more per degree than solids?
2: Discuss the peculiar expansion of water.
3: Explain why a deep lake does not freeze as rapidly as a shallow pond.
4: When a lake freezes, why is the ice formed at its surface?
5: When a liquid is heated in a dilatometer, we can observe a slight depression in the level of the liquid before the apparent expansion takes place. Explain.
6: Distinguish between the coefficients of real and apparent expansion of a liquid. What is the relation between them? Write the equation.

PROBLEMS

1: If the coefficient of absolute expansion of a certain liquid is 0.0014 per centigrade degree, find the value per Fahrenheit degree.
2: An ordinary glass flask of volume 1000 cubic centimeters is full of mercury at 20°C. What volume will overflow when the temperature is raised to 50°C?
3: A tank contains 10 gallons of alcohol at 20°C. Find the increase in volume when the temperature is raised to 30°C.
4: A tank contains 20 cubic feet of alcohol at 90°F. Find the decrease in volume when the temperature drops to 40°F.

5: A steel tank, when completely filled at 20°C, holds 16 gallons of gasoline. If the temperature rises to 30°C, how much gasoline will overflow?

6: A copper tank when completely filled at 0°C will hold 100 gallons of turpentine. If the temperature rises to 25°C, how much turpentine will overflow?

7: Ten gallons of ethyl alcohol are placed in a tank when the temperature is 25°C. What will be the change in volume of this alcohol when the temperature drops to 0°C?

8: A pyrex dilatometer contains exactly 10.000 cubic centimeters of mercury at 20°C. The mercury is then heated to 100°C and the volume reading is 10.144 cc. Find
   a. the apparent expansion of mercury
   b. the coefficient of apparent expansion of mercury
   c. the coefficient of absolute expansion of mercury
   d. the absolute expansion of mercury

9: What is the increase in volume of a gallon of carbon tetrachloride when its temperature is increased from 0°C to 30°C?

10: One cubic centimeter of mercury at 20°C has a mass of 13.55 grams. What will be the mass of 1 cubic centimeter of mercury at 100°C?

11: The bulb and stem of a mercury thermometer contain 0.5 cubic centimeters of mercury at 0°C. What is the distance between the 0°C and 100°C marks if the internal diameter is 0.02 centimeters? The coefficient of apparent expansion of mercury in this thermometer is 0.0015/C°.

CHAPTER 19 EXPANSION OF GASES

We know that gases expand when heated. If a toy balloon filled with air is heated it will expand, and when it is cooled it will contract. We defined heat as a form of energy. When the temperature of a gas is increased its thermal energy is also increased and so the velocity of the molecules becomes greater. This increase in molecular velocity causes the molecules to move farther apart, and so we say that the volume of a given mass of gas increases with a rise in temperature.

In the preceding two chapters, we studied the expansion of solids and liquids. We noticed that different solids and liquids have different coefficients of expansion. Two questions suggest themselves:

1. Have different gases different coefficients of expansion?
2. Exactly how much does a gas expand when heated?

Two centuries ago these questions attracted the attention of French scientists. J. A. C. Charles, J. L. Gay-Lussac, H. V. Regnault, and other scientists showed by experiment that under constant pressure all gases expand equally for equal increases in temperature, except for those near the liquefying temperature of the gas. The increase in volume experienced by a unit volume of a gas when its temperature is raised 1 centigrade degree is called the *coefficient of expansion* of that gas. Charles made a rough estimate of this ratio. Later, Gay-Lussac measured it more accurately and found it to be approximately 1/273 of its volume at 0°C. This principle may be stated as follows: *The volume of a given mass of gas at constant pressure increases for each rise of 1 centigrade degree by 1/273 of the volume it would occupy at 0°C.*

Suppose we consider the volume of a given mass of gas to be one liter at 0°C. For convenience, let us imagine this to be an "ideal" gas which would not liquefy but would continue to contract at the same rate even at very low temperatures.

The following table shows the approximate volumes of this given mass of gas at different temperatures:

Temperature	Volume (under constant pressure)	
100°C	$1l + \frac{100}{273}l$	or $1\frac{100}{273}l$
50°C	$1l + \frac{50}{273}l$	or $1\frac{50}{273}l$
10°C	$1l + \frac{10}{273}l$	or $1\frac{10}{273}l$
1°C	$1l + \frac{1}{273}l$	or $1\frac{1}{273}l$
0°C	$1l$	
−1°C	$1l - \frac{1}{273}l$	or $\frac{272}{273}l$
−2°C	$1l - \frac{2}{273}l$	or $\frac{271}{273}l$
−10°C	$1l - \frac{10}{273}l$	or $\frac{263}{273}l$
−100°C	$1l - \frac{100}{273}l$	or $\frac{173}{273}l$
−200°C	$1l - \frac{200}{273}l$	or $\frac{73}{273}l$
−271°C	$1l - \frac{271}{273}l$	or $\frac{2}{273}l$
−272°C	$1l - \frac{272}{273}l$	or $\frac{1}{273}l$
−273°C	$1l - \frac{273}{273}l$	or $0l$

From the above table we note that the coefficient of volume expansion for gases is 1/273 of the volume at 0°C, or 0.00366/C°.

DETERMINING THE COEFFICIENT OF EXPANSION OF A GAS

The coefficient of expansion of a gas may be determined in the laboratory by the method outlined on the following page.

Take a long piece of capillary tubing of uniform bore (Fig. 19-1) sealed at one end. Heat the tube slightly to expel some of the air and, at the same time, dry the remaining air. Introduce a thread of mercury into the tube, trapping enough of the dry air in the tube to occupy about half the length of the tube at room temperature. Attach the tube to a meter stick and immerse in a tall glass vessel filled with melting ice. When the mercury has come to rest, read the length of the air column. Now replace the ice and water with hot water, stirred so that the temperature is the same throughout. When the mercury has come to rest again, read the length of the air column. Take the temperature reading by inserting a thermometer into the water bath.

During such an experiment, the following observations were made:

Length of air column at 0°C = 35.2 cm.

Length of air column at 80°C = 45.5 cm.

Since the tube is of uniform bore, the volume is proportional to the length of the column. Assume the cross-sectional area of the tube to be 1 sq. cm.

35.2 cc. for a rise of 80 C° expands
45.5 cc. − 35.2 cc. = 10.3 cc.

1 cc. for a rise of 1 C° expands $\dfrac{10.3 \text{ cc.}}{35.2 \text{ cc.} \times 80 \text{ C°}}$

$$= 0.00366/\text{C°} \text{ or } \frac{1}{273 \text{ C°}}$$

In doing this experiment, a number of volume readings should be taken at different temperatures. The coefficients of expansion should be calculated between these temperatures, and then the average coefficient should be found. Why? One should remember that in this experiment no correction has been made for the expansion of the glass.

THE KELVIN OR ABSOLUTE TEMPERATURES

Since the volume of a given mass of gas, for each degree of rise in temperature, increases by 1/273 of its volume at 0°C, it follows that the volume will decrease by 1/273 of its volume at 0°C for each degree fall in temperature.

Fig. 19-1. Apparatus for measuring the expansion of a gas.

Fig. 19-2. A temperature-volume graph (temp. °C).

Let us take another look at the temperature-volume table on page 177. At 0°C the volume of this given mass of gas is 1 liter. If the temperature falls to −1°C the volume is diminished by 1/273 and the gas now occupies a volume of 272/273 of its volume at 0°C, or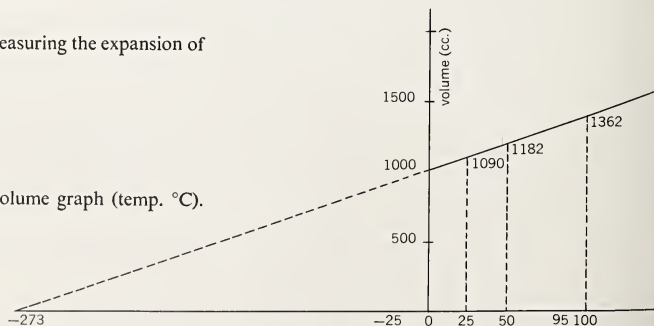 272/273 liter; at −2°C the volume is 271/273 liter; at −100°C the volume is 173/273 liter; and so on. From this line of reasoning the volume would disappear at −273°C (see Fig. 19-2). From the graph, it appears that −273°C is the lowest temperature possible. We know that materials possess thermal energy. Decreasing the temperature results in the removal of thermal energy. At −273°C a substance is completely deprived of heat, and is at the lowest possible temperature. There is no longer any molecular motion and so no spaces between adjacent molecules. We say the substance has reached *absolute zero.*

Lord Kelvin (1824–1907) devised a temperature scale using absolute zero (−273°C) as the zero point. The degrees are the same size as centigrade degrees. This scale is called the *absolute* or the *Kelvin* scale. Fig. 19-3 shows the comparison on the centigrade and Kelvin, or absolute, thermometers. From the diagram we see that an absolute temperature is 273 degrees higher than the corresponding centigrade temperature. To change a centigrade reading to an absolute reading, add 273 degrees to the centigrade reading. In equation form,

$$A = C + 273$$

Let us rewrite the temperature—volume table shown on page 177 in terms of absolute degrees and study the results.

Temperature (°A)	Volume (liters)	$V \div T$
373	373/273	1/273
323	323/273	1/273
283	283/273	1/273
274	274/273	1/273
273	1	1/273
272	272/273	1/273
271	271/273	1/273
263	263/273	1/273
173	173/273	1/273
73	73/273	1/273
2	2/273	1/273
1	1/273	1/273

From this table we see that the volume of a given mass of gas varies directly as the
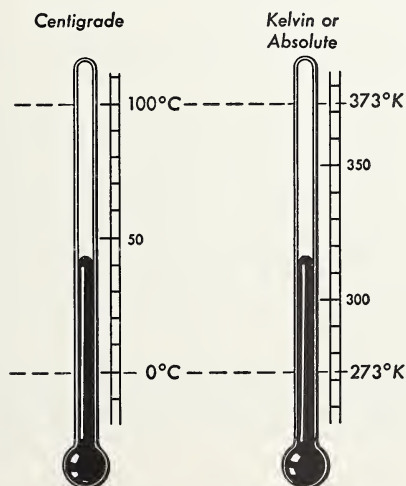

Fig. 19-3. Comparison of centigrade and absolute scales.

absolute temperature. *Charles's law* may be stated as follows:

If the pressure is constant, the volume of a given mass of gas is directly proportional to the absolute temperature.

Note that the quotient obtained when the volume of a given mass of gas is divided by the absolute temperature of the gas is always constant. Why? In equation form, Charles's law may be written as

$$\frac{V}{T} = K \text{ (if pressure is kept constant)}$$

or

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

where T_1 and T_2 are absolute temperatures.

Example 1: *An enclosed quantity of gas occupies 50 cc. at 27°C. What would be its volume at 77°C, the pressure remaining the same?*

$V_1 = 50$ cc. $\qquad V_2 = ?$

$T_1 = 27°C$ or $300°A$ $\quad T_2 = 77°C$ or $350°A$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{50 \text{ cc.}}{300°A} = \frac{V_2}{350°A}$$

$$V_2 = \frac{50 \text{ cc.}}{300°A} \times 350°A = 58.3 \text{ cc.}$$

THE GENERAL GAS LAW: A COMBINATION OF CHARLES'S AND BOYLE'S LAWS

In dealing with gases, it often happens that both temperature and pressure conditions change. To find the new volume that a given mass of gas occupies under such conditions, it is necessary to combine the effects of Charles's law and Boyle's law.

In Chapter 13 we stated Boyle's law in equation form, as

$$\frac{V_1}{P_2} = \frac{V_2}{P_1} \text{ (temperature constant)}$$

and in this chapter we stated Charles's law as

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ (pressure constant)}$$

Combining the two we get a general gas law which covers all cases. In equation form it may be written as

$$\frac{V_1 P_1}{T_1} = \frac{V_2 P_2}{T_2}$$

Example 2: *A given mass of gas occupies a volume of 144 cc. at 0°C and 760 mm. of mercury pressure. Find its volume at 18°C and 770 mm. of mercury pressure.*

$V_1 = 144$ cc. $\qquad V_2 = ?$

$T_1 = 0°C$ or $273°A$ $\quad T_2 = 18°C$ or $291°A$

$P_1 = 760$ mm. $\qquad P_2 = 770$ mm.

$$\frac{V_1 P_1}{T_1} = \frac{V_2 P_2}{T_2}$$

$$\frac{144 \text{ cc.} \times 760 \text{ mm.}}{273°A} = \frac{V_2 \times 770 \text{ mm.}}{291°A}$$

$$V_2 = \frac{144 \text{ cc.} \times 760 \text{ mm.}}{273°A}$$

$$\times \frac{291°A}{770 \text{ mm.}} = 151.5 \text{ cc.}$$

STANDARD TEMPERATURE AND PRESSURE

Often, when dealing with gases, both the temperature and pressure conditions at which they are measured must be noted. These conditions would have to be the same for a comparison of volumes to have significance. We need a standard. The *standard* or *normal* temperature selected is 0°C and the standard or normal pressure is 760 mm. of mercury. This is often abbreviated to *S.T.P.* or *N.T.P.*

Example 3: *The volume of a gas at 120°C and 722 mm. is 22.4 liters. Find its volume at S.T.P.*
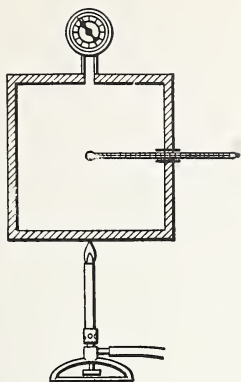
$V_1 = 22.4l$ $\qquad V_2 = ?$

Fig. 19-4. Effect of temperature on the pressure of a gas.

Fig. 19-5. The variation of pressure with temperature at constant volume.

$T_1 = 120°C$ or $393°A$ $T_2 = 273°A$

$P_1 = 722$ mm. $P_2 = 760$ mm.

$$\frac{V_1 P_1}{T_1} = \frac{V_2 P_2}{T_2}$$

$$\frac{22.4l \times 722 \text{ mm.}}{393°A} = \frac{V_2 \times 760 \text{ mm.}}{273°A}$$

$$V_2 = \frac{22.4l \times 722 \text{ mm.}}{393°A} \times \frac{273°A}{760 \text{ mm.}} = 14.8l$$

RELATION BETWEEN PRESSURE AND TEMPERATURE

Suppose we have a given amount of gas in an airtight box which has a pressure gauge and a thermometer attached to it (Fig. 19-4). The volume of the enclosed gas remains constant. If the gas is heated, its pressure increases. If cooled, the pressure decreases. It can be shown by experiment that if the temperature of the gas is increased from 0°C to 1°C, the pressure is increased by 1/273 of its value at 0°C. If the temperature is lowered to −10°C, the pressure is decreased by 10/273 of its value at 0°C. At −100°C the pressure is decreased by 100/273, at −200°C by 200/273, and so on.

By following this line of reasoning we

conclude that the pressure would become zero at −273°C (Fig. 19-5).

At −273°C the molecules have been deprived of all their thermal energies. There is no longer any movement and so no pressure.

The above principle may be summarized as follows:

When the volume is constant, the pressure of a gas is proportional to the absolute temperature.

In equation form,

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Example 4: A given mass of gas exerts a pressure of 76 centimeters of mercury at 27°C. Find the pressure at 57°C, if the volume of the gas remains constant.

$P_1 = 76$ cm. $P_2 = ?$

$T_1 = 27°C$ or $300°A$ $T_2 = 57°C$ or $330°A$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{76 \text{ cm.}}{300°A} = \frac{P_2}{330°A}$$

$$P_2 = \frac{76 \text{ cm.}}{300°A} \times 330°A = 83.6 \text{ cm.}$$

Example 5: A pressure gauge shows that the pressure in an automobile tire is 30 pounds per square inch when the temperature is 27°C. What is the pressure reading when the temperature rises to 45°C? The barometric pressure is 15 pounds weight per square inch.

$P_1 = 30$ lb/in.$^2 + 15$ lb/in.$^2 = 45$ lb/in.^{2}

$$P_2 = ?$$

$$T_1 = 300°A.$$ $$T_2 = 318°A.$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{45 \text{ lb/in.}^2}{300°A} = \frac{P_2}{318°A}$$

$$P_2 = \frac{45 \text{ lb/in.}^2}{300°A} \times 318°A = 47.7 \text{ lb/in.}^2$$

Pressure gauge reading is

$$47.7 \text{ lb/in.}^2 - 15 \text{ lb/in.}^2 = 32.7 \text{ lb/in.}^2$$

THINGS TO REMEMBER

Gases expand when heated.

Under constant pressure all gases expand equally for equal increases in temperature.

Coefficient of expansion of gases is 1/273 of their volume at 0°C.

Absolute zero is the temperature at which all the energy has been removed from a body.

$$0°A = -273°C$$

To change centigrade readings to absolute readings use

$$A = C + 273$$

Charles's law states that the volume of a given mass of gas is directly proportional to the absolute temperature, if the pressure is kept constant.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

S.T.P. or N.T.P. refers to 0°C and 760 mm. of mercury pressure.

When the volume is constant, the pressure of a gas is proportional to the absolute temperature.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

QUESTIONS

1: What is meant by the term "a given mass of gas"?
2: Why do gases expand upon heating?
3: Define (a) "ideal" gas, (b) absolute zero.
4: What is the coefficient of expansion of a gas?
5: If the volume of a given mass of gas is kept constant, why does the pressure increase with an increase in temperature?
6: State Boyle's law in equation form.
7: State Charles's law in equation form.
8: What is the general law of gases?

9: If the absolute temperature of a given mass of gas is doubled while the pressure is kept constant, find the effect on (a) its volume, (b) its mass, (c) its density.

10: Why is it necessary to have standard conditions of temperature and pressure?

PROBLEMS

1: Change the following centigrade temperatures to corresponding absolute temperatures: (a) 0°, (b) 20°, (c) 100°, (d) −200°, (e) −273°.

2: Change the following absolute temperature to corresponding centigrade temperatures: (a) 100°, (b) 0°, (c) 273°, (d) 373°.

3: A mass of gas at 20°C has a volume of 400 cubic centimeters. Pressure remaining constant, find its volume at (a) 0°C, (b) 100°C.

4: Calculate the volumes occupied at the given final pressures by gases whose initial volumes and pressures are listed below. The temperature remains constant in each case.

	Initial Volume	Initial Pressure	Final Pressure
a.	100 cc.	760 mm.	700 mm.
b.	500 cc.	740 mm.	760 mm.
c.	3.50 liters	700 mm.	standard pressure
d.	4.75 cu. ft.	standard pressure	80 cm.
e.	10.0 cu. ft.	30 lb. wt/sq. in.	40 lb. wt/sq. in.
f.	22.4 liters	1 atmosphere	1.5 atmospheres

5: Calculate the volumes which would be occupied at the given final temperatures by gases the initial volumes and temperatures of which are given below. The pressure remains constant in each case.

	Initial Volume	Initial Temperature	Final Temperature
a.	500 cc.	0°C	27°C
b.	4.50 cu. ft.	15°C	100°C
c.	22.4 liters	27°C	0°C
d.	100 cc.	standard temperature	97°C

6: Calculate the volumes which would be occupied at the given final temperatures and pressures by gases the initial volumes and temperatures of which are given below.

	V_1	T_1	P_1	T_2	P_2
a.	22400 cc.	0°C	760 mm.	27°C	700 mm.
b.	300 cc.	0°C	760 mm.	−17°C	740 mm.
c.	22.4l	120°C	900 mm.	0°C	760 mm.
d.	9.75 cu. ft.	−37°C	70 cm.	0°C	76 cm.
e.	27.5 cu. ft.	68°C	45 lb. wt/sq. in.	32°F	15 lb. wt/sq. in.

7: A mass of oxygen occupies 600 cubic centimeters at 27°C and 60 cm. of mercury. Find its volume at N.T.P.

8: A sealed flask of air at atmospheric pressure and 20°C has its temperature lowered to −120°C. Neglecting any volume change in the flask, find the final pressure in the flask.

9: One thousand cubic centimeters of gas are heated from 27°C to 420°C, while at the same time the external pressure is raised from 15 to 45 pounds per square inch. Find the new volume.

10: An automobile tire shows a pressure of 24 pounds weight per square inch above atmospheric pressure at 20°C. Find the change in volume of the air if the tire is punctured. The atmospheric pressure is 15 pounds weight per square inch.

11: The density of a gas is 1.43 grams per liter at S.T.P. Find its density at 27°C and 700 mm. of mercury.

12: A certain mass of gas occupies 35.2 cubic centimeters at 0°C and 48.1 cubic centimeters at 100°C. Find the mean coefficient of expansion between these temperatures.

20 SPECIFIC HEAT AND THERMAL CAPACITY

UNITS OF HEAT

In defining units for measuring thermal energy, water is used as a standard. In the metric system the unit of heat is the *calorie.*

A calorie is the quantity of heat required to raise the temperature of 1 gram of water 1 centigrade degree. It requires 1 calorie of heat to raise the temperature of 1 gram of water through 1 centigrade degree. It requires 60 calories of heat to raise the temperature of 10 grams of water through 6 centigrade degrees. To find the number of calories required to heat water, multiply the number of grams of water by the change in temperature in centigrade degrees. We can state this equation form as follows,

$$H = m(t_2 - t_1)$$

where H is expressed in calories if m represents the mass of water in grams and $(t_2 - t_1)$ the temperature change in centigrade degrees.

Example 1: How many calories of heat are required to raise 1000 grams of water from 20°C to 80°C?

Using $\qquad H = m(t_2 - t_1)$

$$H = 1000(80 - 20)$$
$$H = 60,000$$

∴ the amount of heat required is 60,000 calories.

In the English system, the **British thermal unit** is used (abbreviated B.t.u.).

A British thermal unit is the quantity of heat required to raise the temperature of 1 pound of water 1 Fahrenheit degree.

One B.t.u. equals 252 calories.

THERMAL CAPACITY

If we were to take 10 grams of water, 10 grams of aluminum, and 10 grams of iron at the same temperature and try to raise their temperatures by heating, we would find that water takes the most heat, and iron the least. In fact, it takes 10 calories of heat to raise 10 grams of water 1 centigrade degree; 2.2 calories to raise 10 grams of aluminum 1 centigrade degree; and 1.1 calories to raise 10 grams of iron 1 centigrade degree (Fig. 20-1).

Different materials require varying amounts of heat to undergo the same
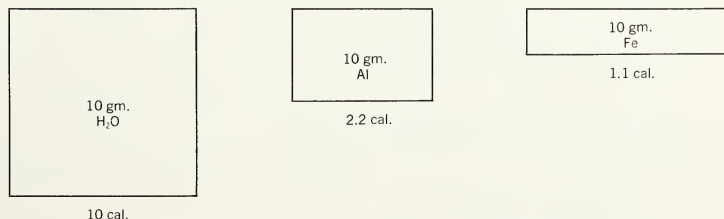

Fig. 20-1. Heat required for temperature change of 1C°.

temperature change, even though the materials have the same mass. We say that such substances differ in *heat capacity*.

The heat capacity of a body is defined as the quantity of heat required to raise its temperature 1 centigrade degree.

SPECIFIC HEAT

Since water is the standard for defining units of heat, it is often necessary to compare the thermal capacity of each substance with that of water. This ratio is called the *specific heat* of the substance.

Specific heat of a substance is defined as the ratio of the thermal capacity of that substance to the thermal capacity of an equal mass of water.

$$\text{Specific heat of substance} = \frac{\text{Thermal capacity of substance}}{\text{Thermal capacity of equal mass of water}}$$

Example 2: If the thermal capacity of 10 grams of aluminum is 2.2 calories, find its specific heat.

$$\text{Specific heat of aluminum} = \frac{2.2 \text{ cal.}}{10 \text{ cal.}} = 0.22$$

Note that in the above definition the specific heat is expressed as a pure number. It should also be noted that the specific heats of substances are numerically equal in both the metric and English systems.

In solving some types of heat problems, it is often necessary to express specific heat in units. This is essential if we are to show the units explicitly and treat them as algebraic expressions. For this purpose we shall redefine *specific heat* of a substance as:

The number of calories required to raise the temperature of 1 gram of the substance 1 centigrade degree

or *The number of B.t.u. required to raise the temperature of 1 pound of the substance 1 Fahrenheit degree.*

In the metric system the specific heat is expressed in calories per gram per centigrade degree, and in the English system in British

thermal units per pound per Fahrenheit degree.

The amount of heat required to raise 1 gram of aluminum through 1 centigrade degree is 0.22 calories, so we say its specific heat is 0.22 cal/gm/C°.

In the English system, the amount of heat required to raise 1 pound of aluminum 1 Fahrenheit degree is 0.22 B.t.u., so we say that the specific heat of aluminum is 0.22 B.t.u./lb/F°.

Table of Specific Heat Values

Substance	Specific Heat Calories per gram per centigrade degree or B.t.u. per pound per Fahrenheit degree
Water	1.00
Alcohol	0.66
Aluminum	0.22
Copper	0.09
Glycerine	0.57
Ice	0.53
Iron	0.11
Lead	0.03
Silver	0.06

HEAT REQUIRED TO RAISE THE TEMPERATURE OF A SUBSTANCE

To find the heat required to raise the temperature of a certain substance, multiply the mass of the material by the temperature change by its specific heat.

$$H = m(t_2 - t_1)S$$

In the metric system, H is expressed in calories if m is in grams, $(t_2 - t_1)$ in centigrade degrees, and S in cal/gm/C°.

In the English system, H is expressed in B.t.u. if m is in pounds, $(t_2 - t_1)$ in Fahrenheit degrees, and S in B.t.u./lb/F°.

Example 3: How much heat is required to raise the temperature of 20 grams of iron from 20°C to 70°C?

$$H = m(t_2 - t_1)S$$

$H = 20$ gm. $\times 50\text{C}° \times 0.11$ cal/gm/C°

$H = 110$ cal.

Example 4: How much heat is required to raise the temperature of 100 pounds of aluminum from $-10°F$ to $100°F$?

$H = m(t_2 - t_1)S$

$H = 100$ lb. $\times 110\text{F}° \times 0.22$ B.t.u./lb/F°

$H = 2420$ B.t.u.

TO FIND THE THERMAL CAPACITY OF A SUBSTANCE

From our definition of a calorie, it follows that the thermal capacity of 10 grams of water is 10 calories. That is, 10 calories of heat are required to raise the temperature of 10 grams of water one centigrade degree.

This can be shown by using

$H = m(t_2 - t_1)S$

$H = 10$ gm. $\times 1\text{C}° \times 1$ cal/gm/C°

$H = 10$ cal.

Example 5: Find the heat capacity of 40 grams of copper.

To raise 1 gm. of copper 1C°, the amount of heat required is 0.09 cal.
To raise 40 gm. of copper 1C°, the amount of heat required is 3.6 cal.
∴ the heat capacity of 40 gm. of copper is 3.6 cal.

Or, using

$H = m(t_2 - t_1)S$

$H = 40$ gm. $\times 1\text{C}° \times 0.09$ cal/gm/C°

$H = 3.6$ cal.

Since, by our definition of heat capacity, the value of $(t_2 - t_1)$ is always 1 degree, we say that *heat capacity = mS.*

If we use units of measurement explicitly in the above equation, the value of heat capacity is expressed in cal/C°.

Heat capacity of 40 gm. of copper =

40 gm. $\times 0.09$ cal/gm/C° = 3.6 cal/C°

WATER EQUIVALENT

In heat measurement, the term *water equivalent* is often used.

The water equivalent of a body is the mass of water which has the same thermal capacity as the body.

Thus, if the thermal capacity of a body is 15.5 calories per centigrade degree, its water equivalent is 15.5 grams, because this is the number of grams of water which has a thermal capacity of 15.5 calories per centigrade degree. Note that the thermal capacity expressed in calories is numerically equal to the water equivalent expressed in grams.

Example 6: If the mass of a piece of iron is 100 grams and its specific heat is 0.11, find its (a) thermal capacity, (b) water equivalent.

a. To raise 1 gm. of iron by 1C° we need 0.11 cal. To raise 100 gm. of iron by 1C° we need 11 cal.
∴ the thermal capacity of iron is *11 cal/C°.*

b. The water equivalent of iron is *11 gm.* since water equivalent is numerically equal to thermal capacity.

Alternate solution

a. Thermal capacity = $m(t_2 - t_1)S$
= 100 gm. $\times 1\text{C}°$
$\times 0.11$ cal/gm/C°
= 11 cal/C°

b. Water equivalent = mS
= 100 gm. $\times 0.11$
= *11 gm.*

Note that in *water equivalent = ms* the thermal capacity ratio definition of specific heat is used, since a pure number is desired to obtain the correct unit explicitly.

PRÉVOST'S LAW OF HEAT EXCHANGE (Pierre Prévost: 1751–1839)

When you pour cold cream into a cup of hot coffee, the cream is heated and the coffee is cooled until both the coffee and the cream

reach the same temperature. Some of the heat energy of the coffee is lost to the cream.

The law of conservation of energy states that energy can be neither created nor lost, simply changed from one form to another. In this case the coffee molecules are slowed down and the cream molecules are speeded up, until both the coffee molecules and the cream molecules have the same average energy of motion.

Heat lost = heat gained

Example 7: A 100-gram piece of iron at a temperature of 100°C was dropped into an aluminum vessel containing 50 grams of water at 20°C. The mass of the aluminum vessel is 30 grams. Find the final temperature.

Let the final temperature be t°C.

Heat lost by iron $= m(t_2 - t_1)S$
$$= 100 \text{ gm.}(100 - t)\text{C}°$$
$$\times 0.11 \text{ cal/gm/C}°$$
$$= (1100 - 11t) \text{ cal.}$$

Heat gained by

a. water $\quad = 50 \text{ gm.}(t - 20)\text{C}°$
$$\times 1 \text{ cal/gm/C}°$$
$$= (50t - 1000) \text{ cal.}$$

b. aluminum vessel $= 30 \text{ gm.}(t - 20)\text{C}°$
$$\times 0.22 \text{ cal/gm/C}°$$
$$= (6.6t - 132) \text{ cal.}$$

But heat lost = heat gained
$$\therefore (1100 - 11t) \text{ cal.} = (50t - 1000) \text{ cal.}$$
$$+ (6.6t - 132) \text{ cal.}$$
$$1100 - 11t = 56.6t - 1132$$
$$67.6t = 2232$$
$$t = 33$$

$\therefore$ the final temperature is *33°C*.

SPECIFIC HEAT OF SOLIDS

The specific heat of a substance may be found by using Prévost's law of heat exchange. A suitable vessel in making heat measurements by this method is called a *calorimeter*. A calorimeter consists of a polished metal vessel supported by a fiber ring within another vessel, with an air space between the two vessels. The metal vessel is polished to minimize absorption or radiation of heat. The air space acts as an insulator. A wooden cover with one hole for a thermometer and one for a mixing spoon is provided (Fig. 20-2).

Suppose we wish to find the specific heat

Fig. 20-2. A calorimeter.

of a block of lead in the laboratory. The procedure is as follows:

1. Determine the substance of the inner vessel of the calorimeter. (Supposing it to be copper, S.H. = 0.09.)
2. Weigh the inner vessel of the calorimeter.
3. Put some water in the calorimeter and weigh again, to obtain the weight of water.

4. Place the inner vessel, with the water in it, into the outer vessel and take the temperature of the water.
5. Weigh the piece of lead.
6. Attach a light string to the lead and heat in boiling water. Take the temperature of the boiling water.
7. Quickly transfer the lead from the boiling water to the calorimeter. (Use the string in doing this.)
8. Stir the cold water and watch the thermometer. As soon as the temperature becomes steady, read the thermometer.

Observations

Mass of lead = 500 gm.
Initial temperature of lead = 100°C
Mass of inner vessel = 80 gm.
Specific heat of inner vessel = 0.09
Mass of water in vessel = 100 gm.
Initial temperature of vessel and water = 15°C
Final temperature of vessel and water = 25°C

Calculations

Let x cal/gm/C° be the specific heat of lead.

Heat lost by lead = 500 gm.$(100 - 25)$C°
$$\times \frac{x \text{ cal.}}{\text{gm/C}°}$$
$$= 37,500x \text{ cal.}$$

Heat gained by

a. water = 100 gm.$(25 - 15)$C° $\times \frac{1 \text{ cal.}}{\text{gm/C}°}$
$$= 1000 \text{ cal.}$$

b. vessel = 80 gm.$(25 - 15)$C° $\times \frac{0.09 \text{ cal.}}{\text{gm/C}°}$
$$= 72 \text{ cal.}$$

But heat lost = heat gained
$$37,500x \text{ cal.} = 1000 \text{ cal.} + 72 \text{ cal.}$$
$$37,500x = 1072$$
$$x = 0.028$$

∴ specific heat of lead is $\dfrac{0.028 \text{ cal.}}{\text{gm/C}°}$

If, in the above calculations, we let x be the specific heat instead of x cal/gm/C°, then

heat lost = 500 gm.$(100 - 25)$C° $\times x$
$$= 37,500x \text{ gm.C}°$$
heat gained is still 1072 cal.
$$37,500x \text{ gm.C}° = 1072 \text{ cal.}$$

$$x = \frac{1072 \text{ cal.}}{37,500 \text{ gm/C}°}$$

$$= \frac{0.028 \text{ cal.}}{\text{gm.C}°}$$

SPECIFIC HEAT OF LIQUIDS

Prévost's law of heat exchange may also be used to measure the specific heat of liquids. The procedure is similar to that described in the previous section. The liquid, of unknown specific heat, is placed in the inner vessel of the calorimeter. A hot block of metal, of known specific heat, is dropped into the liquid. Then the heat given up by the hot metal would be equal to the heat gained by the calorimeter and the liquid.

Suppose the following observations were made:

Mass of lead	= 500 gm.
Specific heat of lead	$= \dfrac{0.03 \text{ cal.}}{\text{gm/C}°}$
Initial temperature of lead	= 100°C
Mass of inner vessel	= 80 gm.
Specific heat of inner vessel	$= \dfrac{0.09 \text{ cal.}}{\text{gm/C}°}$
Mass of alcohol	= 150 gm.
Initial temperature of vessel and alcohol	= 20°C
Final temperature of lead and alcohol	= 30°C

Calculations

Let x cal/gm/C° be the S.H. of alcohol.

Heat lost by lead = 500 gm.$(100 - 30)$C°
$$\times \frac{0.03 \text{ cal.}}{\text{gm/C}°}$$
$$= 1050 \text{ cal.}$$

Heat gained by

a. calorimeter $= 80$ gm.$(30 - 20)$C°

$$\times \frac{0.09 \text{ cal.}}{\text{gm/C°}} = 72 \text{ cal.}$$

b. alcohol $\quad = 150$ gm.$(30 - 20)$C°

$$\times \frac{x \text{ cal.}}{\text{gm/C°}} = 1500x \text{ cal.}$$

Heat gained $=$ heat lost

$1500x$ cal. $+ 72$ cal. $= 1050$ cal.

$1500x$ cal. $= 978$ cal.

$x = 0.65$

∴ specific heat of alcohol is $\dfrac{0.65 \text{ cal.}}{\text{gm/C°}}$

THINGS TO REMEMBER

A *calorie* is the quantity of heat required to raise the temperature of 1 gram of water 1 centigrade degree.

A ***British thermal unit*** is the quantity of heat required to raise the temperature of 1 pound of water 1 Fahrenheit degree: 1 B.t.u. $= 252$ cal.

Heat capacity is the quantity of heat required to raise the temperature of a body 1 centigrade degree.

Specific heat is the quantity of heat required to raise the temperature of 1 gram of a body 1 centigrade degree, or 1 pound of water 1 Fahrenheit degree.

Water equivalent of a body is the mass of water which has the same thermal capacity as the body.

Prévost's law of heat exchange: ***Heat lost = heat gained.***

Heat gained or lost by a substance may be found as follows: $H = m(t_2 - t_1)S$.

QUESTIONS

1: Define a calorie.
2: Define a British thermal unit.
3: What is the relationship between a calorie and a British thermal unit?
4: Define thermal capacity.
5: Why does a bar of steel feel colder in winter than a block of wood, although the temperatures are the same?
6: Define water equivalent.
7: What happens when a hot object is brought in contact with a cold one?
8: How does the specific heat of aluminum compare with that of lead?
9: Cooking utensils often have wooden handles. Why?
10: Why does a large body of water take a longer time to freeze than a small one?
11: What is a calorimeter?
12: How would you measure the specific heat of a metal?

PROBLEMS

1: How many calories of heat are required to raise the temperature of:
   a. 10 grams of iron from 20°C to 50°C?
   b. 20 grams of water from 0°C to 100°C?
   c. 50 grams of iron from 20°C to 900°C?
2: How many B.t.u. are required to raise the temperature of:
   a. 15 pounds of water from 40°F to 150°F?
   b. 20 pounds of lead from 68°F to 218°F?
   c. 10 pounds of aluminum from 20°C to 100°C?
3: Find the thermal capacity of:
   a. 10 grams of water
   b. 150 grams of aluminum
   c. 2.5 pounds of iron.
4: Find the water equivalent of:
   a. a copper calorimeter (mass 120 grams)
   b. a glass rod (mass 20 grams)
   c. an iron vessel (mass 2 pounds)
5: A silver spoon has a mass of 40 grams. How much heat does it absorb if the temperature is raised from 20°C to 80°C?
6: If 100 grams of a metal are heated to 100°C and then dropped in 120 grams of water at 10°C, the resulting temperature is 15°C. Find the specific heat of the metal.
7: A piece of iron, mass 89 grams and temperature 90°C, is placed in 70 grams of water at 10°C. If the resulting temperature is 20°C, find the specific heat of iron.
8: If 100 grams of water at 95°C are poured into 80 grams of water at 20°C, find the resulting temperature. (Neglect the container.)
9: A copper calorimeter the mass of which is 100 grams contains 100 grams of water. How many calories are required to raise the temperature from 22°C to 30°C?
10: An aluminum kettle weighing 2.5 pounds and containing 6 pounds of water at 68°F is heated to 212°F. How many B.t.u. are required?
11: A calorimeter of mass 200 grams contains 50 grams of water at 10°C. Another 100 grams of water at 80°C is added. The final temperature is 50°C. Find the specific heat of the calorimeter.
12: A calorimeter contains 80 grams of a liquid at 10°C. When 100 grams of a metal at 100°C are placed in the liquid, the resulting temperature is 25°C. Specific heat of metal is 0.09. Water equivalent of calorimeter is 10 grams. Calculate the specific heat of the liquid.

CHAPTER 21 LATENT HEAT

CHANGE OF STATE

When a solid is heated, its temperature rises. At a certain temperature, the solid melts and becomes a liquid. This change of state, from a solid to a liquid, is often called *fusion* or *liquefaction*. The temperature at which a body changes from a solid to a liquid under normal atmospheric pressure is called its *melting point*.

When a liquid is cooled, it changes into a solid. This change of state from a liquid to a solid is called *freezing* or *solidification*. The temperature at which solidification occurs under normal pressure is called the *freezing point*. *Crystalline* substances have definite melting or freezing points. *Noncrystalline* substances like paraffin have no definite melting point. When heated, they soften gradually before melting.

Like a solid, when a liquid is heated, its temperature rises, and at a certain temperature the liquid changes to a vapor. This change of state, from a liquid to a vapor or gas, is called *vaporization*. The temperature at which this substance, under normal atmospheric pressure, changes from a liquid to a vapor state is called its *boiling point*.

The most common example of change of state is the freezing and the boiling of water. If a block of ice at −20°C is uniformly heated, its temperature rises slowly until it reaches 0°C. At 0°C the temperature stops rising and the ice begins to melt. The temperature remains constant until all the ice melts. As soon as all the ice has turned to water, the temperature begins to rise again, and continues to rise until it reaches 100°C. At 100°C, boiling sets in and again the temperature stops rising until all the water is changed to steam. When this has been accomplished at 100°C the temperature begins to rise once more. The graph (Fig. 21-1) shows these changes of temperature and changes of state for 1 gram of ice starting at −20°C.

Fig. 21-1. Temperature-heat graph for 1 gram of ice.
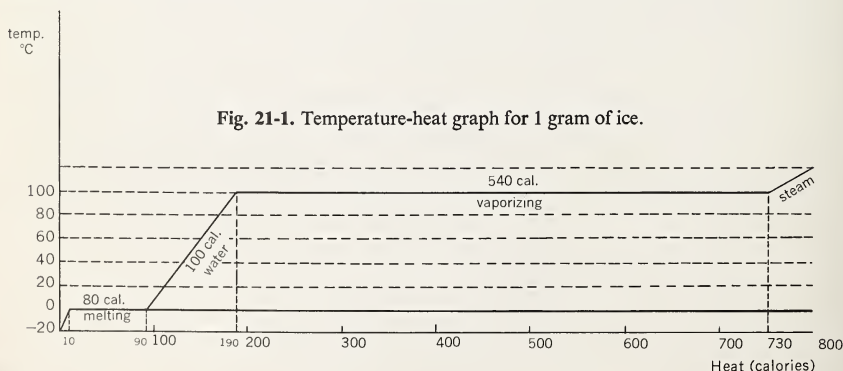
LATENT HEAT OF FUSION

From the temperature-heat graph we can see that ice at 0°C absorbs heat in changing to water at this point. Since there is no change in temperature, the heat absorbed by the ice in melting is called *latent* or hidden. This heat increases the energy of the molecules of the ice. We know that the molecules of ice, or of any solid, vibrate about fixed positions. As their energies are increased the molecules move far enough apart so that they can slip around each other.

Latent heat of fusion is defined as the number of calories required to change 1 gram of a substance from a solid to a liquid state without a change in temperature.

To change 1 gram of ice at 0°C to water at 0°C approximately 80 calories of heat are required. We say that the latent heat of fusion of ice is 80 calories per gram.

In the English system, the *latent heat of fusion is the number of B.t.u. required to change 1 pound of a substance from a solid to a liquid state without a change in temperature.*

The latent heat of fusion of ice is 144 B.t.u. per pound.

MEASURING THE HEAT OF FUSION OF ICE

The heat of fusion of ice may be measured in the laboratory by the method of mixtures. For example, let 100 grams of finely broken ice at 0°C be dropped into 400 grams of water at 40°C contained in a calorimeter the water equivalent of which is 10 grams. The mixture is stirred until all the ice is melted and the temperature reading settles at say 16.5°C.

Calculation

Let x cal/gm. be the latent heat of fusion of ice.

Heat gained by melting ice = 100 gm. $\times$ x cal/gm. = $100x$ cal.

Heat gained by resulting water = 100 gm.(16.5 − 0)C° $\times$ 1 cal/gm/C° = 1650 cal.

Total heat gained = $(100x + 1650)$ cal.

Heat lost by water and calorimeter

$$= 410 \text{ gm.}(40 - 16.5)C° \times 1 \text{ cal/gm/C°}$$
$$= 9635 \text{ cal.}$$

But heat gained = heat lost

$$(100x + 1650) \text{ cal.} = 9635 \text{ cal.}$$
$$100x = 7895$$
$$x = 79.85$$

∴ the latent heat of fusion of ice is *80 cal/gm.*

LATENT HEAT OF VAPORIZATION

From the temperature-heat graph we can see that water at 100°C absorbs heat in changing to steam without a change in temperature. The number of calories of heat required to change 1 gram of water at 100°C to 1 gram of steam at 100°C is known as the latent heat of vaporization of water. The latent heat of vaporization of water is approximately 540 calories per gram.

Latent heat of vaporization may be defined as the number of calories of heat required to change 1 gram of a liquid into vapor without changing its temperature.

In the English system, the *latent heat of vaporization is the number of B.t.u. required to change 1 pound of a liquid into a gas without changing its temperature.*

The latent heat of vaporization of water is 972 B.t.u. per pound.

MEASURING THE HEAT OF VAPORIZATION OF WATER

The heat of vaporization of water may be determined in the laboratory by the following method: Pass steam from a boiler or flask for a few minutes into a quantity of water in a calorimeter. Take the weight and the temperature of the water before and after the steam is passed into it. The difference in weights is equal to the weight of water from the condensation of steam. To prevent water formed by condensation in the delivery tube from entering the calorimeter, a steam trap is used (Fig. 21-2).

Suppose the mass of water in the calorimeter at first is 90 grams; the water
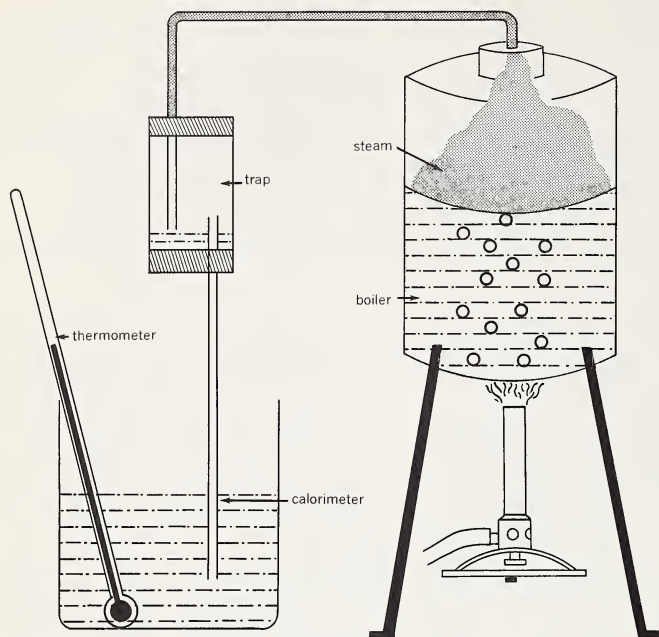
Fig. 21-2.

Finding

the heat

of vaporization

of water.

equivalent of the calorimeter is 10 grams; the mass of steam condensed is 5 grams; the initial temperature of water is 20°C; and the final temperature is 49.5°C.

Calculation

Let x cal/gm. be the heat of vaporization of water.

Heat lost by

a. steam in condensing = 5 gm. × x cal/gm.
$$= 5x \text{ cal.}$$

b. decrease in temp. of 5 gm. of water
$$= 5 \text{ gm.}(100 - 49.5)\text{C}° \times 1 \text{ cal/gm/C}°$$
$$= 252.5 \text{ cal.}$$

Heat gained by water and calorimeter
$$= 100 \text{ gm.}(49.5 - 20)\text{C}° \times 1 \text{ cal/gm/C}°$$
$$= 2950 \text{ cal.}$$

But heat lost = heat gained

$$(5x + 252.5)\text{cal.} = 2950 \text{ cal.}$$
$$5x = 2697.5$$
$$x = 539.5$$

∴ the latent heat of vaporization of water = *540 cal/gm.*

THINGS TO REMEMBER

Fusion is a change of state from a solid to a liquid.

Melting point is the temperature at which a body changes from a solid to a liquid under normal pressure.

Solidification is a change of state from a liquid to a solid.

Vaporization is a change of state from a liquid to a vapor.

Latent heat of fusion is the number of calories required to change 1 gram of a substance

from a solid to a liquid state without a change in temperature. Latent heat of fusion of ice is 80 calories per gram of 144 B.t.u. per pound.

Latent heat of vaporization is the number of calories of heat required to change 1 gram of a liquid into vapor without changing its temperature. Latent heat of vaporization of water is 540 calories per gram of 972 B.t.u. per pound.

QUESTIONS

1: Define "change of state."
2: Does the addition of heat always cause a rise in temperature?
3: What is the difference between a crystalline and a noncrystalline substance?
4: Explain how heat melts a solid.
5: What is meant by heat of fusion?
6: Why does a gram of steam produce a more severe burn than a gram of boiling water?
7: A pan of water may be placed in a vegetable bin to prevent the freezing of vegetables. Explain.
8: Why does the temperature rise during a snowfall?
9: What becomes of heat energy used in melting ice?
10: Explain how evaporation causes a cooling effect.
11: Why is the heat which operates in the melting of ice called "latent heat"?

PROBLEMS

1: How many calories of heat are required to:
   a. Melt 10 grams of ice at 0°C
   b. Change 10 grams of water at 100°C to steam at 100°C
   c. Raise the temperature of 20 grams of ice from −20°C to melting point
   d. Change 10 grams of ice at −20°C to water at 80°C
   e. Change 15 grams of ice at −5°C to steam at 100°C
   f. Change 20 grams of ice at −10°C to steam at 120°C?
2: How much heat is required to vaporize 3.5 pounds of water at 212°F?
3: How much heat is required to change 2 pounds of ice at 32°F to steam at 212°F?
4: How much heat must be removed to change 20 grams of steam at 100°C to ice at 0°C?
5: If 24 grams of steam at 100°C are forced into 500 grams of water at 20°C, find the resulting temperature.
6: In an experiment, 10 grams of steam at 100°C were passed into 200 grams of water at 10°C. Find the latent heat of vaporization of water if the final temperature was 40°C.
7: How many grams of steam at 100°C will be required to raise the temperature of 200 grams of water from 20°C to 50°C?
8: How many grams of steam at 100°C will just melt 20 grams of ice at 0°C?
9: What is the resulting temperature when 200 grams of ice at 0°C are placed in 100 grams of water at 50°C?
10: What mass of water at 100°C must be added to 100 grams of ice in order that all the ice may be melted?
11: A calorimeter of mass 200 grams and specific heat of 0.10 cal/gm/C° contains 400 grams of water at 45°C; 200 grams of ice were melted in it, and the resulting temperature was 5.0°C. Find the heat of fusion of ice.
12: A calorimeter of mass 120 grams and specific heat of 0.1 cal/gm/C° contains 402 grams of water at 6°C. When 23.5 grams of steam at 100°C were passed through it, the resulting temperature was 40°C. Find the latent heat of vaporization of water.

EXPERIMENT 1

22-1. To determine the coefficient of linear expansion (see Fig. 17-5, p. 170).

Apparatus

A coefficient of linear expansion tube with a spherometer, a thermometer, a metal rod, delivery tubes, a meter stick, a steam generator.

Procedure

1. Measure the length of the rod at room temperature with a meter stick.
2. Place the rod in the steam jacket, making sure that it is in contact with the fixed end.
3. Adjust the spherometer so that the central screw just touches the end of the rod.
4. Record the spherometer reading and the room temperature.
5. Loosen the screw of the spherometer.
6. Insert the thermometer in the steam jacket as shown in the diagram.
7. Connect the steam jacket to the generator by using the delivery tube.
8. Pass the steam through the jacket until the mercury stops rising.
9. Adjust the spherometer.
10. Record spherometer and thermometer readings.

Observations

Original length of rod $= l$
Initial temp. of rod $= t_1$
Initial spherometer reading $= S_1$
Final spherometer reading $= S_2$
Final temperature $= t_2$

Calculations

Increase in length of rod $= S_2 - S_1$
Increase in temperature $= t_2 - t_1$
Coefficient of linear expansion $= \dfrac{S_2 - S_1}{l(t_2 - t_1)}$

Note: Show units of measurement explicitly in your calculations.

Percentage Error

Find out what metal the rod is made of. Find the coefficient of linear expansion of this metal. Calculate your percentage error thus:

$$\frac{\text{actual error}}{\text{correct value}} \times 100\% =$$

EXPERIMENT 2

22-2. To determine the coefficient of expansion of a liquid

Apparatus

A dilatometer; a thermometer; a water bath; some glycerine, alcohol, or other liquid.

Procedure

1. Fill the bulb of the dilatometer with the liquid.
2. Place the dilatometer with the liquid in the water bath. Allow to stand for a few minutes.
3. Record the volume of the liquid in the dilatometer and the temperature reading.
4. Change the temperature of the water bath by heating or cooling. Again record the volume of the liquid in the dilatometer and the temperature.

Observations

Temperature	Volume of liquid
1.	
2.	

Calculations

Initial volume of liquid $= V_1$
Initial temperature of liquid $= t_1$
Final volume of liquid $= V_2$
Final temperature of liquid $= t_2$
Coefficient of apparent expansion of the liquid $= \dfrac{(V_2 - V_1)}{V_1(t_2 - t_1)}$

Note: Take a number of volume and temperature readings and find the average of the coefficients.

Coefficient of real expansion of the liquid $=$
$C_A + C_G$

Percentage Error

Coefficient of real expansion of the liquid (from table) $=$
Coefficient (from experiment) $=$
Error $=$

Percentage error $= \dfrac{\text{error}}{\text{correct value}} \times 100\% =$

EXPERIMENT 3

To study the peculiar expansion of water

Apparatus

A dilatometer, a cooling bath, a thermometer.

Procedure

1. Fill the bulb and part of the stem of the dilatometer with water at room temperature.

2. Place the dilatometer and thermometer into the cooling bath (a mixture of ice and water).
3. As the water in the dilatometer cools, keep a record of the volumes at different temperatures.

Observations (make a chart like the one at the foot of the page).

Conclusions

1. What is peculiar about the expansion of water?
2. At what temperature is water at its maximum density?

EXPERIMENT 4

22-3. To determine the expansion of a gas at constant pressure (see Fig. 19-1, p. 178).

Apparatus

About 60 centimeters of capillary tubing sealed at one end, mercury, a meter stick, a centigrade thermometer, a tall glass vessel.

Procedure

1. Heat the tube gently to expel some of the air and dry the air.
2. Introduce a thread of mercury to trap some air in the tube.
3. Attach the glass tube and thermometer to the meter stick as shown in Fig. 22-3.
4. Immerse the tube and thermometer in melting ice. When the mercury has come to rest, take the thermometer reading and the length-of-air column.
5. Replace the ice water with hot water. Stir so that the temperature is the same throughout.
6. Immerse the tube and thermometer in hot water. When the mercury has come to rest, take the thermometer reading and the length-of-air column.

Temperature (°C)	20	15	10	9	8	7	6	5	4	3	2	1	0
Volume (cc.)													

Observations

Temperature reading	Column of air
1.	
2.	

Calculations

Change in temperature $= T_1 - 0$
Change in volume $= V_1 - V_0$
Coefficient of expansion of air $= \dfrac{(V_1 - V_0)}{V_0(T_1 - 0)}$

Note: This experiment should be repeated two or three times and the average of the coefficients calculated.

Conclusion

State the coefficient of expansion of air.

Questions

1. Why is this apparatus imperfect?
2. Why is it advisable to repeat this experiment and take the average of the coefficients?
3. What law does this experiment verify?
4. What is the true coefficient of expansion of air?
5. Calculate your percentage error.

EXPERIMENT 5

To determine the specific heat of a metal

Apparatus

Insulated calorimeter, thermometer, beaker, balance, set of weights, metal block.

Procedure

1. Weigh the block of metal.
2. Heat the metal block in a beaker of water and take the temperature of the water when it boils.
3. Weigh the inner vessel of the calorimeter.
4. Find the specific heat of the calorimeter.
5. Put some cold tap water in the inner vessel of the calorimeter. Weigh again to find the weight of the water.

6. Place the inner vessel in the outer and take the temperature of the water.
7. Quickly transfer the metal block from the boiling water to the calorimeter. Use a light string for this.
8. Stir the water and record the temperature as soon as the mercury stops rising.

Observations

Mass of metal block =
Initial temperature of block =
Mass of inner vessel =
Specific heat of calorimeter =
Mass of cold water =
Initial temperature of cold water and calorimeter is =
Final temperature of water, lead, and calorimeter =

Calculations

Let the specific heat of the metal be x cal/gm/C°.
Heat lost by the metal =
Heat gained by the water =
Heat gained by the calorimeter =
Heat lost = heat gained.
(Solve for x. Use units explicitly and treat as algebraic expressions.)

Percentage Error

Compare your result with a value obtained from authentic tables of heat constants. Calculate the percentage error.

EXPERIMENT 6

To determine the specific heat of a liquid:

Apparatus

Insulated calorimeter; thermometer; beaker; balance; set of weights; a metal block of known specific heat; glycerine, alcohol, or some other liquid.

Procedure

1. Weigh the metal block.
2. Heat the metal block in a beaker of water and take the temperature of the water when it boils.
3. Weigh the inner vessel of the calorimeter.

4. Find the specific heat of the calorimeter.
5. Put some of the liquid into the inner vessel of the calorimeter and weigh again to find the weight of the liquid.
6. Take the temperature of the liquid in the calorimeter.
7. Quickly transfer the metal block from the boiling water to the calorimeter.
8. Stir the liquid and record the temperature as soon as the mercury stops rising.

Observations

Mass of metal block =
Specific heat of metal =
Initial temperature of block =
Mass of calorimeter =
Mass of calorimeter and liquid =
Mass of liquid =
Initial temperature of liquid and calorimeter =
Final temperature of liquid, calorimeter, and metal =
Specific heat of calorimeter =

Calculations

Let the specific heat of the liquid be x cal/gm/C°.
Heat lost by the metal =
Heat gained by the liquid =
Heat gained by the calorimeter =
Heat gained = heat lost
(solve for x)

Percentage Error

Compare your result with a value obtained from tables. Calculate the percentage error.

EXPERIMENT 7

To determine the heat of fusion of ice

Apparatus

Insulated calorimeter, thermometer, finely broken ice.

Procedure

1. Weigh the inner vessel of the calorimeter.
2. Find the specific heat of the calorimeter.
3. Put some warm water in the calorimeter.

4. Weigh the calorimeter with the water to obtain the mass of water.
5. Record the temperature of the water.
6. Dry the finely broken ice by placing in a paper towel.
7. Drop the ice in the calorimeter, cover, and stir.
8. As soon as the mercury in the thermometer stops dropping, record the thermometer reading.

Observation

Mass of calorimeter =
Specific heat of calorimeter =
Mass of water and calorimeter =
Initial temperature of water =
Mass of calorimeter, original water, and water from ice =
Final temperature of water =

Calculations

Let the heat of fusion of ice be x cal/gm.

Heat gained

a. to melt ice =
b. to raise resulting water from 0°C to final temperature =

Heat lost

a. By water falling from initial temperature to final temperature =
b. By calorimeter falling from initial temperature to final temperature =
Heat gained = heat lost
Obtained value =
Percentage error =

Question

Why must the ice be dried before placing in water?

EXPERIMENT 8

22-4. To determine the heat of vaporization of water (see Fig. 21-2, p. 194).

Apparatus

Insulated calorimeter, thermometer, a steam generator, a water trap, a short piece of glass or rubber delivery tube.

Procedure

1. Weigh the inner vessel of the calorimeter.
2. Find the specific heat of the calorimeter.
3. Place some water in the calorimeter.
4. Weigh the calorimeter with the water in it.
5. Find the initial mass of water by subtracting the mass of the calorimeter from the mass of the calorimeter and water.
6. Record the initial temperature of the water.
7. Pass steam for a few minutes into the water in the calorimeter. (Set up the apparatus as shown in Fig. 21-2.)
8. Weigh the calorimeter and water after passing steam.
9. Record the final temperature of the water.
10. Find the mass of water formed by condensation by subtracting the initial mass of the calorimeter and water from the final mass of the calorimeter and water.

Observation

Mass of calorimeter =

Specific heat of calorimeter =
Initial mass of calorimeter and water =
Initial temperature of water =
Final mass of calorimeter and water =
Final temperature of water =

Calculations

Initial mass of water =
Mass of steam condensed to water =

Heat lost

a. By steam in condensing =
b. By lowering resulting water from steam temperature to final temperature

Heat gained

a. By water increasing its temperature from initial temperature to final temperature
b. By calorimeter increasing its temperature from initial temperature to final temperature =
   Heat lost = heat gained

Obtained value

Percentage Error

Question

What is the purpose of the water trap?
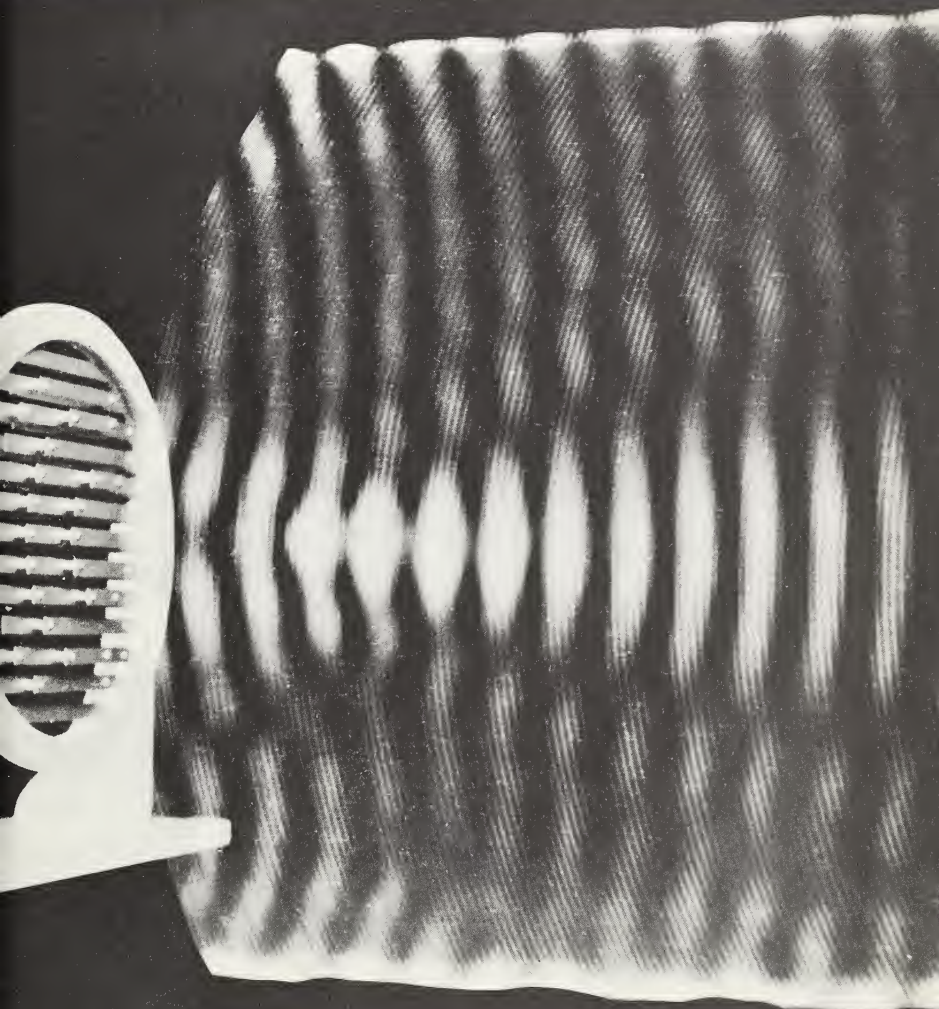
Have you ever thought of what life would be like in a world without sound? For one thing, without sound, we would have difficulty in communicating with each other. While it is true that we have a tendency to take our modern developments in communication for granted, we must remember that scientists spend years of their lives in research laboratories improving our methods of communication.

In this unit, we shall study a few fundamental principles of sound and wave motion.

We shall answer questions such as: What is sound? How is it produced and transmitted? Why are some sounds louder than others? Why are some sounds more pleasing than others? What is the velocity of sound? How can the velocity of sound be calculated in the laboratory?

In studying this unit, let us develop an appreciation for the works of our scientists and an interest keen enough to urge us on to further studies of this subject.

CHAPTER 23   PRODUCTION, PROPAGATION, AND VELOCITY OF SOUND

SOURCES OF SOUND

Pluck a violin or a guitar string and watch it closely. You will notice that sound can be heard only while the string is in vibration. Place a finger on a sounding bell; you will feel a vibration. Strike one of the prongs of a tuning fork with a rubber hammer; you will hear the sound it produces although you may not see the vibration. If you wish to see the vibration, place the tuning fork, prongs down, into a glass of water and you will observe how the prongs spatter the water. Since any matter may vibrate rapidly enough to produce sound, we conclude that *sounds are produced by vibrating matter.*

MEDIUM NECESSARY TO TRANSMIT SOUND

Most of the sounds which we hear travel through air from the vibrating body to our ears. There are other substances also, however, which can convey sound.

Hold your ear against the surface of your desk, and scratch the surface of the desk lightly. You will hear the sound distinctly. Now move your ear away from the surface and scratch lightly again. Can you hear the sound as clearly? If two stones are struck together under water, the sound heard by an ear under water is louder than it would seem if the experiment were performed in air. Thus we see that solids, liquids, and gases all transmit sound, and that solids are better transmitters than liquids and liquids are better transmitters than gases.

Can sound be transmitted in a vacuum? Place a small bell, an alarm clock, or an electric bell connected to flash-light batteries,

under a bell jar, as shown in Fig. 23-1. Pump the air out of the bell jar while the bell is ringing. As the air is pumped out, the sound becomes fainter, until, when most of the air is pumped out, it will not be heard at all. When the air is allowed to enter the bell jar again, the sound becomes louder, enabling us to conclude that sound does not travel through a vacuum. *A material medium is needed for the transmission of sound.*

VELOCITY OF SOUND IN AIR AND OTHER MEDIA

It is common knowledge that sound takes time to travel from one place to another. During a thunder storm, a distant lightning flash is seen seconds before thunder is heard. If we watch a carpenter working at a distance, we see his hammer fall before we hear
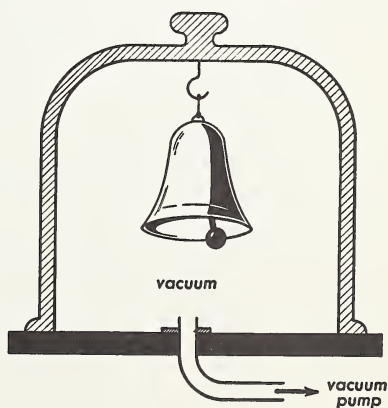

Fig. 23-1. A bell ringing in a vacuum cannot be heard.

the sound. Experiments have shown that *the speed of sound in air at 0°C is 1090 feet per second.* As the temperature rises, the speed of sound increases. *An increase of 1 centigrade degree causes an increase in speed of approximately 2 feet per second.* At 20°C, the speed of sound is about 1130 feet per second. In the metric system, the speed of sound at 0°C is about 332 meters per second, and the increase in speed is about 0.6 meter per centigrade degree.

Experiments have also shown that sound travels faster in liquids than in air. In water sound travels at about 4800 feet per second.

In solids the speed of sound is greater than it is in liquids. In steel, for example, sound travels at about 16,500 feet per second.

Example 1: A thunderclap is heard 5 seconds after the lightning flash was seen. Find the distance of the electrical discharge if the temperature is 15°C.

Velocity of sound at 15°C = 1120 ft/sec.
Time = 5 seconds
Distance = velocity × time
Distance = 1120 ft/sec. × 5 sec. = 5600 ft.

REFLECTION OF SOUND: ECHOES

An echo is a reflected sound wave. A sharp sound made before a wall or a hillside, at a distance of more than 50 feet, strikes the solid and is returned as an echo. The human ear can recognize two separate sounds if they are at least one-tenth of a second apart. If they arrive much closer than this, they blend together into a single sensation. When there are several reflecting surfaces at different distances from the sources, a number of echoes are heard. A succession of echoes is usually heard in the mountains.

Reflected sound travels at the same speed as the original sound. On a day when the temperature is 15°C, the speed of sound is 1120 feet per second. Two seconds elapse before an echo is heard from a cliff 1120 feet distant.

Echoes are useful in sound navigation and ranging, for ships are piloted with the aid of sound echoes. A sound wave is sent out through water by a source mounted on the hull of a ship. The sound wave strikes rock at the bottom, is reflected, and the echo is picked up by a receiver. The time is recorded, and the water depth calculated.

Petroleum engineers use echoes in oil-field explorations. Dynamite explosions are set off in holes drilled into the earth. The vibrations in the earth, called seismic waves, travel in all directions and are reflected to the surface by the deep layers of rock. Receivers stationed at different points pick up and chart these reflected waves. This chart is interpreted to reveal the depth and inclination of the rock formation.

Example 2: The captain of a ship estimates the interval between blowing the horn and hearing the echo to be 2.5 seconds. How far away is the shore (temperature 30°C)?

Velocity of sound at 30°C = 1250 ft/sec.
Total distance sound travels
$$= 1250 \text{ ft/sec.} \times 2.5 \text{ sec.}$$
$$= 3125 \text{ ft.}$$
Distance of shore = 3125 ft. ÷ 2 = 1563 ft.

ACOUSTICS OF AUDITORIUMS

Some materials are good reflectors of sound while others are poor. Usually a hard, smooth surface is a good reflector. If the walls and ceiling of a large auditorium were to be made of hard smooth material, the sound would be reflected many times from one wall to another before it died out. The sound received by the ear would be composed of the original wave and reflected waves of varying strength, and would result in the formation of confused sound. This defect is usually corrected by placing sound-absorbing material over the walls and ceiling. Proper architectural design of the building, as well as of its walls and ceiling, is important in the improvement of its acoustics.

THINGS TO REMEMBER

Sounds are produced by vibrating matter.

A material medium is needed for the transmission of sound.

The speed of sound in air at 0°C is about 1090 feet per second.

An increase of 1 centigrade degree causes an increase in speed of approximately 2 feet per second.

In solids the speed of sound is greater than it is in gases or liquids.

An **echo** is a reflection of sound.

QUESTIONS

1: How are sound waves produced?
2: In what kind of medium can sound travel?
3: Why should the timer of a race stand near the starting gun or watch its smoke?
4: Indians on the Western plains could, by putting the ear to the ground, detect the tramping of buffalo too far away to be seen. Explain.
5: The flash of lightning is seen before thunder is heard. Explain.
6: Knowing the velocity of sound, how could you measure the distance across a lake?
7: Why is it more difficult to hear a speaker outside than inside a building?
8: What is the advantage of using a megaphone?
9: Why are heavy drapes often used on the walls of theaters?
10: Why does the presence of an audience improve the acoustics of a hall?

PROBLEMS

1: Find the speed of sound in air (in feet per second) at the following temperatures: (a) 10°C, (b) 30°C, (c) −20°C, (d) 68°F.
2: Find the speed of sound in air (in meters per second) at the following temperatures: (a) 20°C, (b) 40°C, (c) −10°C.
3: A man shouts, and 3 seconds later hears his echo. How far away is the reflecting body (temperature 68°F)?
4: The sound of thunder was heard 5 seconds after the flash of lightning. How far away was the flash (temperature 20°C)?
5: An echo comes back from the bottom of the ocean 4 seconds after the sound is produced. What is the depth of the ocean?
6: A stone is dropped into a deep well. The sound is heard at the surface $\frac{1}{2}$ second after the stone strikes the water. Find the depth of the water level. (Temperature of air in the well is 5°C.)
7: A man standing 3300 feet from a tall cliff shouts, and hears his echo 6 seconds later. Find the speed of sound.

CHAPTER 24 TRANSMISSION OF SOUND

We have seen that for the hearing of sound a vibrating body and a material medium are necessary. These vibrations are propagated in the medium in the form of waves. Everybody has observed wave motion in matter. No doubt you have watched water waves on a large body of water. It appears that the water particles are moving onward, but we know this is not true. If we watch a log on the surface we observe that, as the waves pass along, the log simply moves up and down but does not move farther from or nearer to us. The motion of the water is handed on but not the water itself. If you were to tie one end of a rope to a post and flip the other end while the rope is fairly taut, you would notice a wave moving along the rope, although you know that the rope itself does not move toward the post. *A wave may be defined as a disturbance which travels without change of form and without the medium moving bodily with it.*

TRANSVERSE VIBRATIONS

Waves produced by jerking the end of a taut rope, or plucking a guitar string are known as *transverse vibrations.* If we examine a transverse vibration more closely, we see that *the particles vibrate at right angles to the path along which the wave travels.*

Fig. 24-1 shows the displacements of a medium transmitting a transverse wave.

Transverse vibrations form *crests* and *troughs.* Particles at the crests (A, B) of the waves are undergoing a maximum displacement upward, those at the bottom of the troughs (C), a maximum displacement
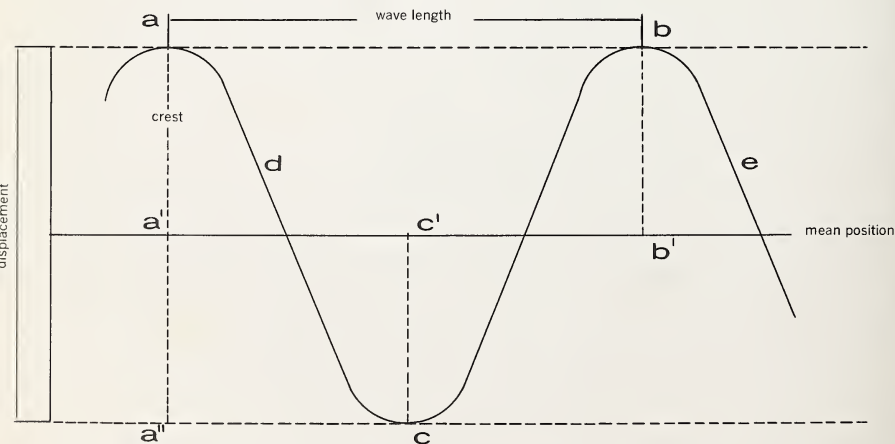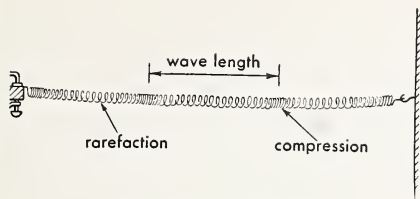

Fig. 24-1. Transverse waves.
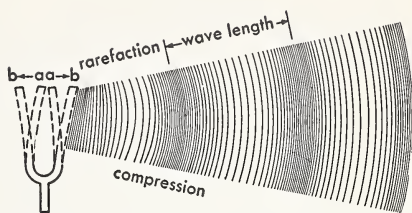
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Fig. 24-2. Longitudinal vibrations.

Fig. 24-3. A vibrating body produces condensations and rarefactions of air particles.

downward. The maximum displacement from the mean position (AA', BB', CC') is known as the *amplitude* of the wave.

Particles at the same distance from their mean positions and moving in the same direction are said to be in the same *phase*. Thus particles A and B are in the same phase. Particles D and E are also in the same phase.

Wave length is the distance between two consecutive particles in the same phase. The distance from crest to crest or trough to trough is a wave length. The wave length represents the distance that the motion has traveled during one complete vibration, or *cycle*. During a cycle, particle A moves to A', to A", back to A' and to A.

The *period* is the time required for one cycle. The *frequency* is the number of complete vibrations or cycles per second.

LONGITUDINAL VIBRATIONS

To illustrate longitudinal vibrations we may use a stretched coil spring (Fig. 24-2). Compress several turns of the spring near one end and release them quickly. This compression travels to the opposite end and returns by reflection. The particles of the spring do not vibrate at right angles to the direction of the wave motion, as in the case of transverse waves, but vibrate to and fro along the path which the wave travels. In longitudinal waves the compressed regions are known as *condensations,* and the stretched regions as *rarefactions.*

SOUND WAVES ARE LONGITUDINAL VIBRATIONS

We know that a vibrating fork produces sound. As the prong of a tuning fork moves from its original position A to the new position B, the air particles immediately to the right of it are compressed, while the particles immediately to the left are rarefied. As the prong moves back from B to A, this process is reversed. As the prong continues to vibrate, a series of condensations and rarefactions of air particles is formed as shown in Fig. 24-3. This series of rarefactions and condensations is known as a longitudinal *train of waves.*

We must remember that when a condensation moves forward, there is very little forward movement of air particles. Each compression simply transfers its energy to the air particles ahead of it. This transfer of energy can be illustrated as shown in Fig. 24-4. If we roll one ball up to the others, it will be stopped by collision with the end ball, the impulse being transferred through each ball in turn. The end ball will roll away but the others will remain in their original position.

Fig. 24-4. Experiment illustrating the law of conservation of momentum.

RELATION BETWEEN VELOCITY, WAVE LENGTH, AND FREQUENCY

While a vibrating body makes one complete vibration or cycle, the wave motion travels one wave length. If we know the wave length produced by a vibrating body and the frequency or the number of complete vibrations the body makes per second, we can find the velocity of any wave motion by multiplying the frequency by the wave length. Thus,

if the wave length $= l$
and the frequency $= n$
then the velocity of wave motion $= nl$
or $V = nl$

Example 1: If the frequency of a tuning fork is 256 and the length of the waves produced is 4.3 feet, calculate the velocity of sound.

$l = 4.3$ ft.
$n = 256$/sec.
$V = nl$
$V = 256$/sec. $\times$ 4.3 ft. $= 1100$ ft/sec.

Example 2: The frequency of a tuning fork is 384 cycles per second. If the velocity of sound in air is 1100 feet per second, find the wave length produced by the tuning fork.

$n = 384$/sec.
$V = 1100$ ft/sec.
$l = ?$
$l = V/n$
$l = \dfrac{1100 \text{ ft/sec.}}{384 \text{/sec.}} = 2.9$ ft.

STANDING WAVES

If two students hold the opposite ends of a rope and each student jerks his hand upward at the same instant starting two pulses along the rope, these pulses will meet at the midpoint causing a great displacement there. If the ends are moved downward at the same instant, the two downward pulses will meet at midpoint. Continued vibrations can be timed so that the rope vibrates up and down in one segment (Fig. 24-5). This is called a *standing wave* because it does not seem to move. If one student moves his end of the rope upward, while the other moves his end downward at the same instant and by the same amount, the two will counteract each other at the midpoint (B), which will be a point of zero motion. This point of zero motion is called a *node*. Maximum motions occur in regions D and D', so that the rope vibrates in two segments. These regions of maximum motion are called *antinodes*. By vibrating the ends more rapidly, any number of nodes and antinodes may be produced. Because the nodes and antinodes do not move along the rope, the resultant waves are called standing waves.

Standing waves may be produced in the laboratory by attaching one end of a light cord to one prong of a large tuning fork and passing the other end over a pulley and attaching it to a pan of weights (Fig. 24-6). By adjusting the length and tension of the cord, standing waves may be produced and studied.

Standing waves play an important part in the production of musical tones. The tone of lowest frequency emitted when the string vibrates in one segment is called the *fundamental*. The others are *overtones*. Their frequencies are 2, 3, 4, or more times that of the fundamental tone. The strings of musical instruments vibrate in a complex manner, giving out a blend of different frequencies.

Fig. 24-5. A string may vibrate in one or more segments.

BEATS

When two sources of sound having slightly different frequencies are sounded together, the two sets of waves will interfere with each other. When the crests of one coincide with the crests of the other, reinforcement occurs. When the crests of one meet with the troughs of the other, the waves tend to neutralize each other. When the longitudinal waves of sound neutralize each other, no sound is heard. When they reinforce each other, the sound is louder. Each recurrence of maximum loudness is called a *beat*.
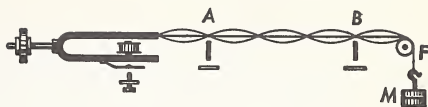
The number of beats per second due to two simple tones is equal to the difference between the two frequencies of the waves.

Thus, if the frequencies of the sound waves which act together are 40 and 50 vibrations per second, 10 beats per second would be produced.

Fig. 24-6. Standing waves.

THINGS TO REMEMBER

In *transverse vibrations* the particles vibrate at right angles to the path along which the wave travels.

Wave length is the distance between two consecutive particles in the same phase.

Period is the time required for one cycle.

Frequency is the number of complete vibrations or cycles per second.

Longitudinal vibrations is the motion of condensations and rarefactions along the path which the wave travels.

Sound waves are longitudinal vibrations.

Relation between velocity, wave length, and frequency may be expressed as

$$V = nl$$

The fundamental is the tone of lowest frequency emitted when the body vibrates in one segment.

Beat is a recurrence of maximum loudness.

The number of beats per second due to two simple tones is equal to the difference between the two frequencies of the waves.

QUESTIONS

1: How is sound transmitted?
2: Distinguish between transverse and longitudinal waves.
3: Explain the following terms: (a) wave length, (b) amplitude, (c) frequency, (d) phase, (e) period, (f) condensations and rarefactions, (g) standing waves, (h) nodes and antinodes, (i) fundamental tones and overtones, (j) beats.
4: How do we know that sound waves of different frequencies travel with the same velocity?
5: What is the relation between frequency, wave length, and the velocity of a wave?
6: Are the outdoor sound waves in air for a given pitch longer in winter or in summer?
7: How can two musical instruments be accurately tuned by listening for beats?

PROBLEMS

1: If the crests of a wave are 2 inches apart, and 6 crests pass by a given point every second, find the velocity of the wave motion.

2: A tuning fork having a frequency of 320 vibrations per second produces air waves 3.56 feet in length.
   a. What is the speed of sound in air?
   b. What is the temperature?

3: A tuning fork which vibrates 275 times per second is used to produce sound when the temperature of the air is 5°C. Find the length of the wave produced.

4: A tuning fork vibrates 256 times per second. What is the wave length of sound produced when the temperature is 30°C?

5: A string vibrating at 264 times per second sends out waves 4.2 feet long. Find the velocity of the sound waves.

6: The length of a sound wave is 175 centimeters when the temperature is 30°C. What is the frequency?

7: A rope 24 feet long vibrates in 4 segments when shaken at the rate of 2 times per second. Find the velocity of the wave motion.

8: A string vibrating 256 times per second is producing sound at the same time as another vibrating at 264 times per second. What is the frequency of the beats?

9: A tuning fork has a frequency of 440 vibrations per second. If another tuning fork of slightly lower pitch is sounded, 5 beats per second are heard. What is the frequency of the second tuning fork?

10: A string has too high a frequency to be in tune with another having a frequency of 512 vibrations per second. If 4 beats per second are heard, what is the frequency of the first string?

Experience has taught us to recognize and describe different sounds. We may describe a thunderclap as very loud; a siren, shrill or high pitched; a whisper, soft and low; and a singer's voice, one of pleasing quality. Each sound has its distinct characteristics. Sounds differ from one another in pitch, loudness, intensity, and quality.

PITCH

Properties of a sound wave which determine pitch may be studied by using a siren disk (Fig. 25-1). The siren disk has a number of rows of equally spaced holes. The outer row has the most holes, while the inner row has the least holes. While the disk is rotated, a stream of air is directed against each row of holes in succession. The stream of air is cut off regularly, forcing air vibrations which produce sound. If the disk is turned more rapidly, the pitch of the sound rises. The pitch falls as rotation slows down. From these experiments we conclude that *the pitch of a sound depends upon the frequency of its vibrations.*

air blast
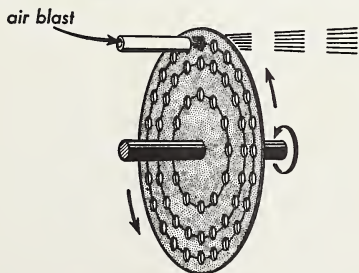
Fig. 25-1. The siren disk.

AUDIBLE FREQUENCIES

If the frequency of a vibration is too low we cannot hear it. For most persons a frequency of at least 20 cycles per second is necessary to affect the auditory nerves. This frequency represents the *lower limit of audibility.* Vibrations may be too rapid to affect the auditory nerves. *The upper limit of audibility is about 20,000 cycles per second.*

Sound waves which vibrate fewer than 20 times per second are known as *infrasonic,* below the range of hearing. Sounds above the 20,000 cycles per second are *ultrasonic.* A dog whistle is an example of ultrasonic waves.

INTENSITY AND LOUDNESS

If you pluck a guitar string gently, a faint sound is produced. If you pluck it hard, the sound produced may be intense and loud. If you watch the vibrations of the string closely, you will notice that the amplitude of the waves is small when the string is plucked gently and the sound is faint. If you pluck the string vigorously, the amplitude of the waves is greater and the sound increases in intensity. A large tuning fork produces stronger condensations and rarefactions than a small one. The sound produced has a greater intensity. *The intensity of a sound depends on the amplitude of the waves and the area of the vibrating body.* The intensity of a tuning fork may be increased by touching a table top with it. Why?

The *loudness* of a sound wave depends upon three factors (1) intensity, (2) distance, and (3) medium. Sounds which are intense do not always seem loud to us. We know

that sound waves are given off in all directions from a vibrating body. The sound waves are spread over greater areas as they move farther away from the source. The areas of these concentric waves are proportional to the squares of their distances from the source. We say that *the loudness of a sound is inversely proportional to the square of the distance from the source.* In a previous chapter it was shown that iron and water are better transmitters of sound than air. *Loudness depends upon the medium through which the sound travels.*

QUALITY

The property which distinguishes one sound from another of the same pitch and loudness is called *quality.* If you play a note on a piano and the same note on a violin, making both notes equal in loudness, the two notes can be easily distinguished even though they have the same pitch and loudness. The difference between them is in the form of waves. *The quality of sound depends upon the number of overtones present in the sound wave.*

RADAR SPEEDOMETER

Science is a progressive and cumulative subject. What we learn in one area of study may be profitably used by alert minds in many other fields. Have you ever noticed how the sound of a rapidly approaching train whistle seems to increase in pitch? Can you explain it? If you live in a city, you may find yourself caught in a so-called "speed trap." The police department knows it as a *radar speedmeter,* and it operates on the same principle as the moving sound situation described above, only it makes use of ultrasonic waves instead of sound waves. The principle here is known as the Doppler effect, discovered first by Professor Doppler in Germany in 1837, using sound waves (Fig. 25-2). Here is how it operates.

The radar speedmeter is a radio transmitter-receiver, operating on the theory that a radio wave, when reflected from a moving target, undergoes a change in wave length (and pitch) proportional to the speed of the target. To simplify this, we will assume that
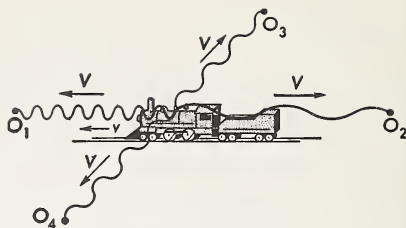

Fig. 25-2. The Doppler effect. The pitch of a whistle on a fast-moving train sounds higher to an observer in front of the train, lower to an observer in back, and normal to observers off at the sides.

the radio wave we are transmitting is 12 inches long.

Transmitter: $12''$ $\xrightarrow{}$ $12''$ $\xrightarrow{}$ $12''$ $\xrightarrow{}$ $12''$ $\xrightarrow{}$ $12''$ $\xrightarrow{}$

If this wave hits a stationary target it will be reflected back to the transmission point with exactly the same wave length.

Transmitter:

$12''$ $12''$ $12''$ $12''$ $12''$
$\overleftrightarrow{12''}$ $\overleftrightarrow{12''}$ $\overleftrightarrow{12''}$ $\overleftrightarrow{12''}$ $\overleftrightarrow{12''}$ I (fixed target)

If, however, while that wave is being reflected, the target moves one inch closer, the wave must then be shortened by one inch when it returns to the transmitter.

Transmitter:

$12''$ $12''$ $12''$ $12''$ $12''$
$\overleftrightarrow{11''}$ $\overleftrightarrow{11''}$ $\overleftrightarrow{11''}$ $\overleftrightarrow{11''}$ $\overleftrightarrow{11''}$ II (target moves one inch)

It can be seen, then, that if the lengths of the transmitted and reflected waves are known, we can determine by their difference, how far the target moved during the time it took the radio wave to reflect from it; that is, the *speed of the target.* When the reflected wave returns, shortened by the movement of the target, its wave length is determined and the difference is used to calculate the speed and show it instantly on a visual scale. Since radio waves travel at the speed of light, the entire operation is instantaneous.

THINGS TO REMEMBER

The *pitch* of a sound depends upon the frequency of its vibration.

The *intensity* of a sound depends on the amplitude of the waves and the area of the vibrating body.

The *loudness* of a sound depends upon intensity, distance, and the medium.

QUESTIONS

1: How is the pitch of sound determined?
2: What characteristics determine the intensity of sound?
3: What three factors determine loudness?
4: What property of sound enables us to distinguish one musical instrument from another?
5: Why is the whistle sound of a rapidly approaching train heard at higher pitch than when the train is standing still?
6: What happens to the pitch of the whistle as the train moves away? Explain.
7: Often a dog perks up his ears and gets excited although his master hears no sound. Explain.
8: What are the limits of human audibility?
9: Why does the pitch of the sound of a saw change?
10: Why is it difficult for one speaking in the open air to be heard? Will a megaphone help?
11: Deaf children listen to music by resting their hands on a piano. Explain.
12: Why does a small child have a high-pitched voice?

CHAPTER 26 RESONANCE: SYMPATHETIC VIBRATIONS

When one child pushes another one seated on a swing, he finds it easier if he makes the frequency of his impulses equal to that of the swing. A man standing on a spring-board can cause great vibrations in the board by jumping up and down at the same rate as the board is vibrating.

Two tuning forks having the same frequency are mounted on individual boxes with open ends. One tuning fork may be placed at one end of the room and the other at the other end of the room, with the open ends of the boxes facing each other as in Fig. 26-1. If we set one of the tuning forks in vibration by striking, and then, after a few seconds, stop the vibration by touching the prongs, we shall note that the other tuning fork is vibrating weakly. What caused the second tuning fork to vibrate? The first tuning fork produced condensations and rarefactions of sound waves in the air. These condensations and rarefactions acted on the second fork in regular succession causing it to vibrate sympathetically. If a violin and a piano are in tune, the A string of the violin may be heard vibrating softly if the A note on the piano is struck. These are examples of *sympathetic vibrations* or *resonance.*

Sympathetic vibrations occur when one body sets another into vibration, the period of the two being the same, or one being a multiple of the other.

RESONANCE IN CLOSED TUBES

Sympathetic vibrations may be demonstrated by holding a vibrating tuning fork near the open end of a tube immersed in a tall jar of water (Fig. 26-2). By moving the tube up and down we find that, at a certain depth, the sound is greatly intensified. This is because the periods of vibration of the fork and of the air column are equal and therefore reinforce each other.

When an air column that is put in vibration by a tuning fork has the same period of vibration as the tuning fork, the air column is said to be in *resonance* with the fork.
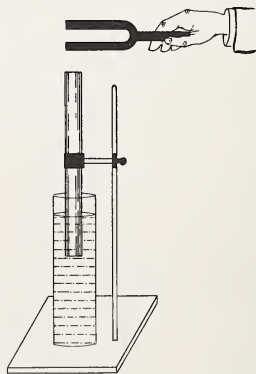

Fig. 26-2. Air column in resonance with a tuning fork in a closed tube.

Fig. 26-1. Resonance between two matched tuning forks.

DETERMINING THE VELOCITY OF SOUND BY RESONANCE

Suppose the lower prong of the tuning fork goes from a to b (Fig. 26-3). It is sending a condensation down the tube. This condensation goes to the bottom of the tube and is reflected upward by the water. For resonance, the tube must be of such a length as to produce the following: The condensation created by the movement from a to b must travel down the tube and be reflected back to a in time to reinforce the condensation set up as the prong moves back from b to a and from b upward. The motion from a to b is a quarter-vibration, producing $\frac{1}{4}$ of a wave length. Therefore, the distance down the closed tube and back again is equal to $\frac{1}{2}$ of a wave length. During a complete vibration the waves travel *four* times the length of the air column in a closed tube. Thus, the length of the air column in a closed tube is $\frac{1}{4}$ of a wave length of sound produced by the fork. In equation form,

$$L_w = 4l_r$$

where L_w is the wave length and l_r is the length of the resonant air column.

It should be pointed out, however, that the diameter of a resonance tube has a slight effect on the wave length. More accurately, one-quarter of the wave length is equal to the length of the resonance tube plus 0.4 of the diameter of the tube or $L_w = 4(l_r + 0.4d)$.

If we know the frequency of the fork, and the length of the resonance column, we can find the velocity of sound.

Example 1: A tuning fork vibrates 256 times per second. A closed-tube air column 13 inches long is in resonance with it. Find the velocity of sound.

Wave length $= 4 \times 13$ in. $= 52$ in. or 4.3 ft.
Velocity of sound $= 256$ cycles/sec. $\times$ 4.3 ft/cycle $= 1100$ ft/sec.

RESONANCE IN OPEN TUBES

Open-tube resonance may be demonstrated by using two tubes, one slipping closely over the other. The tubes should be about 18 inches in length and about $1\frac{1}{2}$ inches in diameter. A vibrating fork is held above the open tube, and the length of the tube is varied by sliding one tube inside the other (Fig. 26-4). At a certain length, the sound is greatly intensified. By measuring the length of the tube it is found that, for the same tuning fork, the open tube is twice as long as the closed tube. Why? In the case of a closed tube, condensations and rarefactions were reflected from the closed end of the tube as condensations and rarefactions. In the open tube the condensation
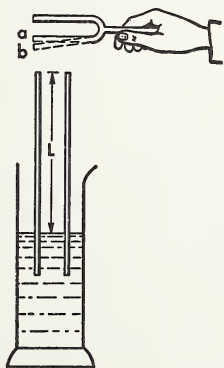

Fig. 26-3. Resonance of an air column in a closed tube.
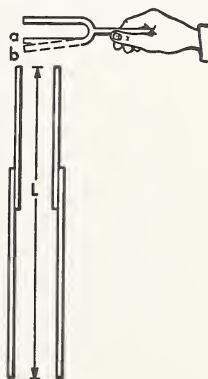

Fig. 26-4. Resonance of an air column in an open tube.

passes out of the open end and a rarefaction is reflected back. For resonance, the prong makes a complete vibration, while the sound wave travels twice the length of the tube. Thus we say that the wave length is approximately twice the length of the air column in an open tube.

$$L_w = 2l_r$$

or more accurately, $L_w = 2(l_r + 0.8d)$.

Example 2: (a) What is the wave length of the sound produced by a tuning fork which resonates with an open tube 40 centimeters long? (b) Find the velocity of the sound waves produced if the frequency of the fork is 412 cycles per second.

Wave length = 2×40 cm. = 80 cm. or 0.8 m.
Velocity of sound wave = 412 cycles/sec. $\times$ 0.8 m/cycle = 330 m/sec.

OVERTONES IN ORGAN PIPES

The most familiar application of the vibrations of air columns is in organ pipes. The length of the air column determines its frequency.

Organ pipes are of two kinds, open and closed (Fig. 26-5). When emitting its fundamental note a closed pipe has a node at the closed end and an antinode at the open end. The length of the air column is equal to the distance from the node to the antinode or $\frac{1}{4}$ of the wave length (Fig. 26-5: A). The same tube can be made to generate overtones of higher frequency than the fundamental by increasing the strength of the air current. When the first overtone is emitted, there are two nodes and two antinodes, and the wave length is one-third as great as in the case of the fundamental (B). Note that in the fundamental $\frac{1}{4}$ of the wave length is standing in the tube and $\frac{3}{4}$ of the wave length is standing in the first overtone. When the second overtone is produced there are $1\frac{1}{4}$ wave lengths standing in the tube and the wave length is one-fifth that of the fundamental (C).

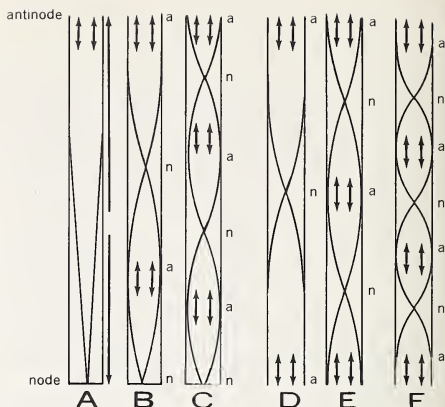The wave lengths of the various overtones produced by closed tubes are 1/3, 1/5,

Fig. 26-5. Standing waves in tubes.

1/7, 1/9, etc., that of the fundamental. The overtones produced by a closed pipe have frequencies of 3, 5, 7, 9, etc., times that of the fundamental.

When emitting its fundamental note, an open tube has a node in the middle and an antinode at each end (D). One-half standing wave length in an open tube produces the fundamental. With a stronger air current, the first overtone is produced with three antinodes, one at the middle and one at each end. Two nodes occur halfway between the antinode as shown in E. In this case the wave length is one-half and the frequency is twice that of the fundamental. The first overtone produced by an open tube has one wave length in the air column. The second overtone produced by an open tube has $1\frac{1}{2}$ wave lengths in the air column (F). The wave length is one-third and the frequency three times that of the fundamental. The open tube forms overtones with frequencies 2, 3, 4, 5, etc., times that of the fundamental. Thus we see that the open tube is richer in overtones than the closed tube. In an open tube all overtones can be produced. In a closed tube only the overtones having a frequency which is an odd multiple of that of the fundamental can be produced.

THINGS TO REMEMBER

Resonance occurs when one body sets another into vibration.

Velocity of sound may be determined by resonance.

The length of the air column in a closed tube is approximately $\frac{1}{4}$ of the wave length of sound produced by the resonating body.

The length of the air column in an open tube is approximately $\frac{1}{2}$ of the wave length of the sound produced by the resonating body.

The overtones produced by a closed tube have frequencies of 3, 5, 7, 9, etc., times that of the fundamental.

The overtones produced by an open tube have frequencies of 2, 3, 4, 5, etc., times that of the fundamental.

QUESTIONS

1: What is a sympathetic vibration?
2: A singer can force a violin string to vibrate and produce the same note. Explain.
3: What is a fundamental tone?
4: What are overtones?
5: How does the frequency of the first overtone compare with that of the fundamental?
6: How does the frequency of the fourth overtone compare with that of the fundamental?
7: How does the frequency of the fifth overtone compare with that of the second overtone?
8: When does resonance occur?
9: What is the relation between the length of a closed resonance tube and the wave length of sound?
10: How does the length of an open tube compare with the wave length of the sound it reinforces?
11: In filling a tall narrow vessel with water one can tell by the sound when it is nearly full. Explain.
12: How can resonance be used to measure the velocity of sound?

PROBLEMS

1: What must be the length of a closed tube to be in resonance with a sound whose wave length is 2 feet?
2: What is the wave length of the fundamental produced by an open pipe 4 feet long?
3: What should be the resonance length of an open tube for a fork whose frequency is 384 vibrations per second if the velocity of sound is 1152 feet per second?
4: An open tube 14 inches long resonates with a fork when the velocity of sound is 1120 feet per second. Find the frequency of the fork.
5: If a closed organ pipe 4 feet long is played, what is the wave length of the fundamental?
6: If an open organ pipe 12 feet long is played, what is the wave length of the fundamental?
7: An open organ pipe produces a fundamental with a frequency of 1760 vibrations per second when the temperature is 25°C. Find the length of the pipe.

8: The frequency of the fundamental tone of an organ pipe is 256 vibrations per second. What are the frequencies of the first and second overtones?

9: A closed tube 12 inches long is in resonance with a tuning fork of 256 vibrations per second. Find (a) the velocity of sound, (b) the temperature of the air.

10: An open tube 18 inches long is in resonance with a tuning fork of 384 vibrations per second. Find (a) the velocity of sound, (b) the temperature of the air.

CHAPTER 27 LAWS OF STRETCHED STRINGS

If we examine a violin we find that it has four strings mounted on a resonance box. The resonance box intensifies the sound. Each of the four strings produces a different note because of its diameter, length, density, and tension. Strings of larger diameter and greater density are used to produce notes of low pitch. The pitch of a note is determined by the frequency of the vibrating string. We know that the frequency of a string may be increased by increasing the tension of the string and by decreasing its length. The tension of a string is altered by turning a key to which the string is attached. The violinist uses his fingers to change the length of the vibrating string. *The frequency of a vibrating string depends on its length, tension, diameter, and density.*

The vibrations of strings may be studied by means of a sonometer (Fig. 27-2). The strings are fastened to steel pins near one end of the resonance box; stretched across two fixed bridges, one at each end of the box; and attached to weights at the other end of the box. The tension of each string can be altered by changing the weight at its end. A movable bridge is used to vary the length in vibration. The vibrations are produced by a bow or by plucking.

LAW OF LENGTHS

Remove the movable bridge and pluck the string. If the string is not too taut, it vibrates as a whole and produces its fundamental note. Place the bridge under the middle of the string and set half of the string in vibration by plucking. The pitch of the note is much higher. As you move the bridge back and forth, lengthening and shortening the length in vibration, the pitch of the note produced is lowered and raised accordingly.

If we remove the bridge and tune a string with a 256-cycle tuning fork and then replace the bridge and move it back and forth until the length in vibration is in tune with a 512-cycle fork, we will find that the length in vibration is exactly half of the original length.

We conclude that, *the frequency of a string varies inversely as its length.*

In equation form,
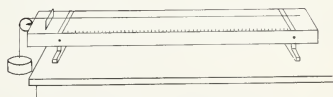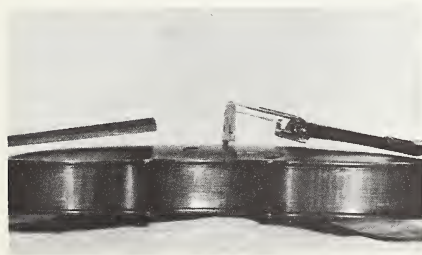
$$\frac{n_1}{n_2} = \frac{l_2}{l_1}$$

Fig. 27-1. The law of stretched strings is illustrated by the violin.

Fig. 27-2. A sonometer.

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Example 1: The frequency of the D note is 288; that of E is 320. If the D string of a violin is 18 inches long, how much must it be shortened to produce the E note?

Since the frequency is inversely proportional to the length, the length of the string producing the E note is 288/320 of 18 inches or 16.2 inches. Therefore, the D string must be shortened by 1.8 inches.

LAW OF TENSIONS

If we alter the tension of a string by changing weights, we soon discover that an increase in the tension produces an increase in the frequency. Let us adjust the tensions of two strings on a sonometer to obtain the same note. That is, the frequencies of the two strings are the same. Now place the movable bridge under the middle of one of the strings. From the law of lengths, we know that the frequency of the vibrating section of the string is twice that of the full-length string. Now keep adding weights to the string of original length until it produces the same note as the string of half length.

If the original tension was 6 pounds weight, the new tension will be 24 pounds weight. That is, in order to double the frequency, the tension of string must be multiplied by four. In order to obtain three times the frequency, we must multiply the tension nine times, and so on.

From the above, we conclude that *the frequency of a vibrating string varies directly as the square root of the tension.*

In equation form,

$$\frac{n_1}{n_2} = \frac{\sqrt{t_1}}{\sqrt{t_2}}$$

Example 2: A string under a tension of 25 pounds weight has a frequency of 50 vibrations per second. Find its frequency under a tension of 9 pounds weight.

Using $\quad \dfrac{n_1}{n_2} = \dfrac{\sqrt{t_1}}{\sqrt{t_2}}$

$$\frac{n_1}{50} = \frac{\sqrt{9}}{\sqrt{25}}$$

$$n_1 = \frac{\sqrt{9}}{\sqrt{25}} \times 50 = \frac{3}{5} \times 50 = 30$$

∴ the frequency is 30 vibrations per second.

LAW OF DIAMETERS

If we use two strings of the same material, the same length, and the same tension, but one having twice the diameter of the other, we find that the frequency of the thicker string is only one-half that of the thinner. If the diameter is three times as great, the frequency is reduced to one-third, and so on.

From the above we conclude that, *the frequency of a vibrating string varies inversely as the diameter of the string.*

In equation form,

$$\frac{n_1}{n_2} = \frac{d_2}{d_1}$$

Example 3: The frequency of a steel wire the diameter of which is 0.6 millimeters is 100 vibrations per second. Find the frequency of a steel wire of the same length and under the same tension if its diameter is 0.2 millimeters.

Using $\quad \dfrac{n_1}{n_2} = \dfrac{d_2}{d_1}$

$$\frac{n_1}{100} = \frac{0.6}{0.2}$$

$$n_1 = \frac{0.6}{0.2} \times 100 = 300$$

∴ the frequency is 300 vibrations per second.

LAW OF DENSITIES

The heavier the string, the slower it vibrates. If we use two strings of the same length, the same diameter, and subjected to the same tension, but with the density of one four times as great as the other, we find that

the frequency of the denser string is only one-half that of the lighter string. If the density is nine times as great, the frequency is reduced to one-third, and so on.

From the above, we conclude that *the frequency of a vibrating string varies inversely as the square root of the density.*

In equation form,

$$\frac{n_1}{n_2} = \frac{\sqrt{D_2}}{\sqrt{D_1}}$$

Example 4: The density of platinum is 21.4 grams per cubic centimeter. The density of silver is 10.5 grams per cubic centimeter. Two strings, one of platinum and one of silver, of the same length, the same diameter, and under the same tension are set in vibration. Compare their frequencies.

Since the density of platinum is approxi-

mately twice as great as that of silver, the frequency of platinum is $1/\sqrt{2}$ or 0.7 times as great as the frequency of silver.

Or, using $\quad \dfrac{n_1}{n_2} = \dfrac{\sqrt{D_2}}{\sqrt{D_1}}$

we get $\quad \dfrac{\text{frequency of platinum wire}}{\text{frequency of silver wire}}$

$$= \frac{\sqrt{10.5}}{\sqrt{21.37}} = \sqrt{\tfrac{1}{2}} = 0.7$$

∴ the frequency of the platinum wire is 0.7 times the frequency of the silver wire.

The four laws of vibrations in strings may be summarized as follows:

The frequency of a vibrating string varies inversely as the length, inversely as the diameter, directly as the square root of the tension, and inversely as the square root of the density.

THINGS TO REMEMBER

The frequency of a vibrating string depends on its length, tension, diameter, and density.

Law of lengths: $\qquad \dfrac{n_1}{n_2} = \dfrac{l_2}{l_1}$

Law of tensions: $\qquad \dfrac{n_1}{n_2} = \dfrac{\sqrt{t_1}}{\sqrt{t_2}}$

Law of diameters: $\qquad \dfrac{n_1}{n_2} = \dfrac{d_2}{d_1}$

Law of densities: $\qquad \dfrac{n_1}{n_2} = \dfrac{\sqrt{D_2}}{\sqrt{D_1}}$

QUESTIONS

1: How do the strings of a piano illustrate all the laws of strings?
2: Which law of strings is used in tuning a violin or a guitar?
3: Which law is used in playing a violin or a guitar?
4: As the movable bridge of a sonometer is moved back and forth, why does the pitch of the note vary?
5: How is the density of the D string of a violin increased?
6: The shorter the string, the higher the pitch. Explain, giving examples from musical instruments.

1 : The frequency of a vibrating string is 400 vibrations per second. What is the frequency if the string is made three-fourths as long?

2 : A string 100 centimeters long has a frequency of 200 vibrations per second. Find the frequency of the note produced by 80 centimeters of the same string if the tension remains the same.

3 : A string has a frequency of 512 vibrations per second when the tension is 16 pounds weight. Find the frequency when the tension is increased to 25 pounds weight.

4 : If the tension of a string producing 256 vibrations per second is 25 pounds weight, find the tension required to produce 320 vibrations per second.

5 : A string of diameter 0.8 millimeters produces 420 vibrations per second. Find the frequency produced by a string of the same material, same length, and subjected to the same tension, if its diameter is 0.2 millimeters.

6 : Two strings of the same length, diameter, and tension weigh 2.5 grams and 10.0 grams respectively. If the frequency of the lighter string is 440 vibrations per second, find the frequency of the heavier string.

7 : Two strings, one of nylon and one of steel, have the same length, same diameter, and are subjected to the same tension. If the frequency of the nylon string is 100 vibrations per second, find the frequency of the steel string. (The density of nylon is 1.1 grams per cubic centimeter; density of steel is 7.7 grams per cubic centimeter.)

8 : One string is 25 centimeters long and 0.5 millimeters in diameter. Another string, made of the same material and under the same tension, is 100 centimeters long and 0.25 millimeters in diameter. If the frequency of the first string is 320 vibrations per second, find the frequency of the second.

9 : When a string 16 inches long is stretched with a force of 49 pounds weight, its frequency is 420 vibrations per second. If the string is shortened to 12 inches and the stretching force is increased to 100 pounds weight, find its frequency.

10 : One wire is twice as long as another, and its tension is 16 times as great. If they are made of the same material and are the same diameter, compare their frequencies.

11 : Two strings have equal lengths, diameters, and tensions but the density of the second is $2\frac{1}{4}$ times that of the first. If the frequency of the first is 120 vibrations per second, find the frequency of the second.

12 : A string 100 centimeters long has a frequency of 320 vibrations per second under a tension of 16 pounds weight. Find the frequency produced by 80 centimeters of the same string under a tension of 9 pounds weight.

13 : A string has a frequency of 100 vibrations per second. If the length, tension, diameter, and density are each multiplied by four, find the frequency.

CHAPTER 28 EXPERIMENTS ON SOUND

EXPERIMENT 1

To show that sound requires a material medium for its transmission

Apparatus

A bell jar, a vacuum pump (complete with base), an electric bell or an alarm clock.

Procedure

1. Place the electric bell or clock inside the bell jar on top of the pump base.
2. Connect the electric bell to a battery. If you are using an alarm clock, make sure the alarm is turned on before placing the clock inside the bell jar.
3. Evacuate the jar with pump.
4. Stop the pump and let air into the bell jar again.

Observations

1. Did you hear the sound when the bell was full of air?
2. What were the changes in loudness as the air was being removed?
3. What were the changes in loudness as the air was returned into the jar?

Conclusion

EXPERIMENT 2

To show standing waves in a stretched chord (see Fig. 24-6, p. 209).

Apparatus

A large tuning fork, a light flexible silk cord, two rigid stands and clamps, a pulley, a pan, a set of weights.

Procedure

1. Attach the tuning fork in a horizontal position to a rigid stand.
2. Attach one end of the silk thread to one prong of the fork.
3. Pass the silk cord over a pulley and attach the other end to a pan of weights.
4. Vary the tension of the cord by adding weights until standing waves are formed. You may find it easier to produce standing waves by adjusting the length of the cord as well as the tension.

Observation

Conclusion

EXPERIMENT 3

To show what determines the pitch of sound (see Fig. 25-1, p. 211).

Apparatus

A disk siren or toothed wheels on a rotating machine, a blow pipe.

Procedure

1. If compressed air is available in your laboratory, connect the blow pipe to the compressed-air tank by using rubber tubing.
2. Send a current of air through the regularly spaced holes near the circumference of the rotating disk.
3. Send a current of air through the middle circle of holes and then through the innermost circle.
4. Vary the speed of the rotating disk and note the changes in pitch.

If compressed air is not available, blow through blow pipe.

Observation

Conclusion

EXPERIMENT 4

a. To demonstrate resonance in a closed tube (see Fig. 26-2, p. 214).
b. To calculate the velocity of sound in air.

Apparatus

A tube about 2 inches in diameter and 18 inches long, a vessel of water, a tuning fork ($n = 256$ is preferable).

Procedure

1. Place one end of the tube in the vessel of water.
2. Hold the vibrating fork over the open end of the tube.
3. Move the tube up and down, until you find a position at which the sound is intensified.
4. Measure the air column in resonance with the fork.

Observations

The frequency of the tuning fork =
The length of the air column =
The room temperature =

Calculation

Frequency = n
Wave length ($4L_r$) = l
Velocity = nl
nl is the velocity of sound at room temperature. Find the velocity at 0°C.

Conclusion

Note: If you use a fork of high frequency on a long tube, an overtone will be produced instead of the fundamental. For above calculations the shortest length in resonance must be found.

EXPERIMENT 5

a. To demonstrate resonance in an open tube.
b. To calculate the velocity of sound in air (see Fig. 26-4, p. 215).

Apparatus

Two tubes about $1\frac{1}{2}$ inches in diameter, one slipping closely over the other, a tuning fork of $n = 256$.

Procedure

1. Vibrate the fork and hold it close to one end of the tube.
2. Vary the length until the sound is intensified.
3. Measure the length of the air column in resonance with the tube.

Observations

The frequency of the fork =
The length of the air column =
The room temperature =

Calculation

Frequency = n
Wave length ($2L_r$) = l
Velocity = nl
Find the velocity of sound at 0°C.

Conclusion

EXPERIMENT 6

To show the production of sympathetic vibrations (see Fig. 26-1).

Apparatus

Two tuning forks mounted on resonance boxes. (The forks must have the same frequency.)

Procedure

1. Place the two tuning forks with the open ends of their resonance boxes facing each other a short distance apart.
2. Vibrate one of them by striking with a mallet.
3. After it has been vibrating for a few seconds, stop the vibration.

Observation

After stopping the vibration of the fork, can you still hear the sound? Where is it coming from?

Conclusion

EXPERIMENT 7

To study the production of beats.

Apparatus

Two tuning forks of the same frequency mounted on resonance boxes, some soft modeling wax or plasticine.

Procedure

1. Stick a piece of wax of approximately the same mass on each prong of one fork.
2. Place the two forks with the open ends of their resonance boxes facing each other a short distance apart.
3. Vibrate the two forks.

Observation

Listen to the throbbing sound caused by a recurrence of maximum loudness. Count the number of beats produced in 30 seconds.

Conclusion

Calculate the number of beats produced per second.
Calculate the frequency of the fork with the wax on each prong.
What causes beats?

EXPERIMENT 8

To show that the frequency of a string varies inversely as its length (see Fig. 27-2, p. 219).

Apparatus

A sonometer with a movable bridge, some weights, two tuning forks having frequencies of 256 and 512.

Procedure

1. Remove the bridge and tune the string

to the 256-cycle fork by altering the tension.
2. Place the bridge under the string and move it back and forth until you find a note in unison with the 512-cycle fork. Do not vary the tension.
3. Measure the length of the string producing the same note as the 512-cycle fork.

Observation

Frequency of fork	Length of string	Frequency × length
$n_1 = 256$	$l_1 =$	
$n_2 = 512$	$l_2 =$	

Note: Use forks of as many different frequencies as are available.

Conclusion

Use the information obtained to show that

$$\frac{n_1}{n_2} = \frac{l_2}{l_1}$$

EXPERIMENT 9

To show that the frequency varies directly as the square root of the tension.

Apparatus

A sonometer with a movable bridge, two strings, some weights.

Procedure

1. Place a suitable weight on string A.
2. Alter the tension of string B until it produces the same note as A. Record the tension.
3. Keep the tension of string A constant for the remainder of the experiment.
4. Place the movable bridge under the middle of A to double the frequency.
5. Keep increasing the tension on B until it is in tune with A. Record the tension.

6. Triple the frequency of string A by moving the bridge to a point one-third of the length of the string.
7. Keep increasing the tension on B until it is in tune with A (one third of string A in vibration). Record the tension.

Observation

Frequency	Tension	Square root of tension
original		
2 × original		
3 × original		

Conclusion

EXPERIMENT 10

To show that the frequency of a string varies inversely as its diameter.

Apparatus

A sonometer, two strings of the same material but one having twice the diameter of the other, some weights.

Procedure

1. Subject the two strings to the same tension.
2. Place the movable bridge at the midpoint of the string the diameter of which is twice as great as the first string.
3. Vibrate the thinner string and half of the thicker string. Listen to the sounds. Are they in unison?
4. If the diameter of one string is three times as great as the first, use the movable bridge to vibrate one-third of its length and listen to the sounds produced.

Observation

Conclusion

Make use of the law of lengths to draw your conclusion.

EXPERIMENT 11

To show that the frequency of a string varies inversely as the square root of its density.

Apparatus

A sonometer, a nylon cord (a fishing line will do) and a steel cord of the same diameter, some weights.

Procedure

1. Subject the two cords to the same tension.
2. Move the bridge under the steel cord back and forth until the steel cord vibrates in unison with the nylon cord.
3. Measure the length of the steel cord in vibration and compare it with the length of nylon cord in vibration.

Observation

Conclusion

Note: The density of nylon is 1.1 gm/cc.

The density of steel is 7.7 gm/cc.

Since the steel is 7.0 times as dense as nylon, the frequency of nylon should be $\sqrt{7.0}$ or 2.6 times as great as the frequency of steel. If this is true, then the length of steel cord vibrating in unison with the nylon cord should be approximately 1/2.6 or 5/13 of the length of nylon (using the law of lengths).

See if the ratio you obtained is approximately 1/2.6 or 0.38

APPENDIX A FORMULAE

1. Mass-density: $D = m/V$

   D = density; m = mass; V = volume

2. Weight-density: $D = W/V$

   D = density; W = weight; V = volume

3. Liquid pressure: $p = hd$

   p = pressure; h = depth; d = weight-density

4. Downward force exerted by a liquid

   (a) $F = Ap$

   F = force; A = area; p = pressure

   (b) $F = Ahd$

   F = force; A = area; h = depth; d = weight-density

5. Sidewise force exerted by a liquid

   (a) $F = Ap/2$

   F = force; $p/2$ = average pressure

   (b) $F = Ahd/2$

   F = force; h = depth; d = weight-density

6. Total pressure (liquid plus atmosphere)

   $$P_T = P_L + P_A$$

   P_T = total pressure; P_L = liquid pressure;

   P_A = atmospheric pressure

7. Force ratio

   $$F.R. = R/F$$

   $F.R.$ = force ratio; R = resistance; F = force

8. Velocity ratio

   $$V.R. = Fd/Rd$$

   $V.R.$ = velocity ratio; Fd = distance force moves;

   Rd = distance resistance moves

9. Efficiency

$$E = F.R./V.R.$$

E = efficiency; $F.R.$ = force ratio;

$V.R.$ = velocity ratio

10. Boyle's law

$$V_1P_1 = V_2P_2$$

V_1 = initial volume; P_1 = initial pressure;

V_2 = final volume; P_2 = final pressure.

11. Charles's law

$$V_1/T_1 = V_2/T_2$$

V_1 = initial volume; T_1 = initial temperature in absolute degrees;

V_2 = final volume; T_2 = final temperature in absolute degrees.

12. Gas laws (combination of Boyle's and Charles's laws)

$$V_1P_1/T_1 = V_2P_2/T_2$$

(as in 10 and 11)

13. Centigrade and Fahrenheit scales

$$C = 5/9(F - 32) \quad \text{or} \quad F = 9/5C + 32$$

C = Centigrade reading;

F = Fahrenheit reading.

14. Linear expansion of solids

$$\Delta l = l(t_2 - t_1)\alpha$$

Δl = linear expansion; l = original length;

t_1 = initial temperature; t_2 = final temperature;

α = coefficient of linear expansion.

15. Coefficient of absolute expansion

$$C_r = C_a + C_c$$

C_r = coefficient of real or absolute expansion.

C_a = coefficient of apparent expansion.

C_c = coefficient of volume expansion of the container.

16. Real or absolute expansion of a liquid

$$\Delta V = V(t_2 - t_1)C_r$$

ΔV = real increase in volume; V = original volume;

t_1 = original temperature; t_2 = final temperature;

C_r = coefficient of absolute or real expansion.

17. Apparent expansion of a liquid

$$\Delta V = V(t_2 - t_1)C_a$$

ΔV = apparent increase in volume; t_1 = initial temperature;

t_2 = final temperature; C_a = coefficient of apparent expansion.

18. Absolute and centigrade scales
$$A = C + 273$$
A = absolute reading; $\quad$ C = centigrade reading.

19. Heat given off or taken in
$$H = m(t_2 - t_1)S$$
H is expressed in calories, if m is expressed in grams, t_1 and t_2 in centigrade degrees and S is the specific heat expressed in cal/gm.C°.

20. Heat capacity
$$\text{Heat capacity} = ms$$
m = mass; $\quad$ s = specific heat.

21. Water equivalent
$$W.e. = ms$$
$W.e.$ = water equivalent; $\quad$ m = mass; $\quad$ s = specific heat.

22. Velocity of sound
$$V = nl$$
V = velocity of sound; $\quad$ n = frequency; $\quad$ l = wave length.

23. Resonance in closed tube
$$L_W = 4l_r$$
L_W = wave length; $\quad$ l_r = length of resonant air column.

24. Resonance in open tube
$$L_W = 2l_r$$
L_W = wave length; $\quad$ l_r = length of resonant air column.

25. Laws of stretched strings

    a. Law of lengths
    $$n_1/n_2 = l_2/l_1$$
    b. Law of tensions
    $$n_1/n_2 = \sqrt{t_1}/\sqrt{t_2}$$
    c. Law of diameters
    $$n_1/n_2 = d_2/d_1$$
    d. Law of densities
    $$n_1/n_2 = \sqrt{D_2}/\sqrt{D_1}$$

    n_1 and n_2 represent frequencies
    l_1 and l_2 = lengths
    t_1 and t_2 = tensions
    d_1 and d_2 = diameters
    D_1 and D_2 = densities

APPENDIX B FORMULAE

Table 1. *Useful Values and Relationships*

12 in. = 1 ft.	1 ft.^{3} of water weighs 62.4 lb. wt.
3 ft. = 1 yd.	1 liter of air at S.T.P. weighs 1.29 gm. wt.
5280 ft. = 1 mi.	1 B.t.u. = 252 calories
1760 yds. = 1 mi.	Speed of sound = 1090 ft/sec. at 0°C.
6080 ft. = 1 nautical mile	Circumference of a circle = $2\pi r$
144 in.^{2} = 1 ft.^{2}	Area of circle = πr^2
9 ft.^{2} = 1 yd.^{2}	Surface of a sphere = $4\pi r^2$
1728 in.^{3} = 1 ft.^{3}	Volume of a sphere = $4/3\pi r^3$
27 ft.^{3} = 1 yd.^{3}	Standard temperature = 0°C
231 in.^{3} = 1 gal.	Standard pressure = 760 mm. of mercury
60 mi/hr. = 88 ft/sec.	Latent heat of vaporization of water 540 cal/gm. or
70000 grains = 1 lb.	972 B.t.u/lb.
16 oz. = 1 lb.	Latent heat of fusion of ice 80 cal/gm. or 144 B.t.u/lb.
2000 lb. = 1 ton	

Table 2. *Metric-English Equivalents*

1 in. = 2.5400 cm.	1 cm.^{2} = 0.1550 in.^{2}
1 ft. = 30.480 cm.	1 in.^{3} = 16.3872 cm.^{3}
1 yd. = 91.440 cm.	1 cm.^{3} = 0.0610 in.^{3}
1 mi. = 1609.4 m.	1 grain = 0.06480 gm.
1 mi. = 1.6094 km.	1 oz. = 28.3495 gm.
1 mm. = 0.03937 in.	1 lb. = 453.592 gm.
1 cm. = 0.3937 in.	1 lb. = 0.4536 kg.
1 m. = 39.37 in.	1 gm. = 15.4324 grains
1 m. = 3.2808 ft.	1 gm. = 0.03527 oz.
1 m. = 1.0936 yd.	1 gm. = 0.002205 lb.
1 km. = 0.621 miles	1 kg. = 2.2046 lb.
1 in.^{2} = 6.4516 cm.^{2}	1 liter = 0.88 (imperial) qt.

Table 3. *Specific Gravity of Solids*

Aluminum	2.7	Human body	1.07
Brass	8.2–8.7	Ice	0.917
Bronze	8.8	Iron, steel	7.6–7.8
Copper	8.9	Limestone	2.7
Cork	0.24	Silver	10.5
Diamond	3.53	Tin	7.3
Glass, crown	2.5	Tungsten	19.3
Gold	19.3	Zinc	7.1
Graphite	2.25		

Table 4. *Specific Gravity of Liquids*
(room temperature)

Alcohol, ethyl	0.789	Glycerine	1.26
Alcohol, methyl	0.793	Hydrochloric acid	1.20
Carbon disulphide	1.29	Kerosene	0.82
Carbon tetrachloride	1.60	Mercury	13.56
Chloroform	1.50	Milk	1.029
Ether	0.74	Water, sea	1.025
Gasoline	0.66–0.69		

Table 5. *Density of Water at Various Temperatures*

°C	gm/cm.³	°C	gm/cm.³	°C	gm/cm.³
0	0.99987	15	0.99913	60	0.98324
1	0.99993	20	0.99823	65	0.98059
2	0.99997	25	0.99707	70	0.97781
3	0.99999	30	0.99567	75	0.97489
4	1.00000	35	0.99406	80	0.97183
5	0.99999	40	0.99224	85	0.96865
6	0.99997	45	0.99025	90	0.96534
8	0.99988	50	0.98807	95	0.96192
10	0.99973	55	0.98573	100	0.95838

Table 6. *Velocity of Sound in Various Media*
(approximate)

material	ft/sec.	material	ft/sec.
Air	1,090	Hydrogen	4,165
Alcohol	3,890	Iron	16,500
Aluminum	16,740	Maple, along grain .	13,470
Brass	11,480	Pine, along grain . .	10,900
Copper	11,670	Steel	16,500
Glass	16,500	Water	4,794

Table 7. *Coefficient of Linear Expansion*
(expansion per unit length per centigrade degree)

Aluminum	0.000023	Lead	0.000029
Brass	0.000019	Platinum	0.000009
Copper	0.000017	Pyrex	0.0000036
Glass (plate)	0.000009	Quartz (fused) . .	0.000004
Gold	0.000013	Silver	0.000017
Ice	0.000051	Steel	0.000013
Iron	0.000012	Tin	0.000027

Table 8. *Coefficient of Volume Expansion*
(increase in volume per unit volume per centigrade degree)

Alcohol (ethyl)	0.0011	Glass	0.000025
Carbon tetrachloride . . .	0.00124	Glycerine	0.00018
Ether	0.0017	Petroleum	0.00096
Gases (at 0°C under		Turpentine	0.00097
constant pressure) . . .	0.0037	Mercury	0.00018
Gasoline	0.00096		

Table 9. *Specific Heat*
(calories per gram per centigrade degree)

Alcohol	0.66	Iron	0.11
Aluminum	0.22	Lead	0.031
Copper	0.09	Mercury	0.033
Glycerine	0.57	Paraffin	0.7
Gold	0.031	Petroleum	0.51
Glass	0.20	Silver	0.06
Ice	0.53		

GLOSSARY

Absolute zero – temperature ($-273°$C) at which all energy has been completely removed.

Acoustics – properties of a room that determine how well sound is heard.

Adhesion – the attraction between unlike molecules.

Antinode – region of maximum motion.

Atmosphere – the ocean of air enveloping the earth.

Barometer – an instrument used to measure atmospheric pressure.

Beat – a recurrence of maximum loudness.

Bernoulli's principle – whenever the velocity of a horizontally moving stream of fluid increases owing to a constriction, the pressure decreases.

Boiling – the process of changing from liquid state to the vapor state. This process takes place throughout the whole body of the liquid.

Boiling point – temperature at which a liquid changes to a vapor under standard pressure.

Boyle's law – the volume of a given mass of gas varies inversely as the pressure, if the temperature remains unchanged.

British thermal unit – the quantity of heat required to raise the temperature of 1 pound of water 1 Fahrenheit degree.

Calorie – the quantity of heat required to raise the temperature of 1 gram of water 1 centigrade degree.

Calorimeter – an insulated cup used to measure the amount of heat.

Capillarity – the rise or fall of a liquid level in capillary tubes above or below the level outside the tube.

Carburetor – a device used to provide the explosion chambers of an internal combustion engine with the proper mixture of gasoline and air.

Charles's law – the volume of a given mass of gas varies directly as the absolute temperature, if the pressure remains the same.

Coefficient of area expansion – the increase in unit area per degree.

Coefficient of linear expansion – the increase in length experienced by a rod of unit length when its temperature is raised 1 degree.

Coefficient of volume expansion – the increase in unit volume per degree. Often referred to as the coefficient of absolute expansion.

Cohesion – the attraction between like molecules.

Condensation – a compression of particles in wave motion.

Crystalline – having a definite melting point.

Dilatometer – a device used to measure the apparent expansion of liquids.

Echo – a reflection of sound.

Energy – capacity for doing work.

Evaporation – the process by which molecules escape from a liquid.

Freezing point – temperature at which a liquid changes to a solid.

Frequency – the number of complete vibrations or cycles per second.

Fundamental – tone of lowest frequency emitted when a string vibrates in one segment.

Fusion – a change from solid to liquid state.

Force – a push or a pull.

Force ratio – the weight supported over the force applied. Often referred to as the real mechanical advantage.

Heat – energy of molecular motion.

Heat capacity – the amount of heat required to raise a body 1 centigrade degree.

Intensity – auditory sensation which depends upon the amplitude of the waves and the area of the vibrating body.

Ionosphere – the region above the stratosphere which contains electrically charged particles.

Law of densities – the frequency of a vibrating string varies inversely as the square root of the density.

Law of diameters – the frequency of a vibrating string varies inversely as the diameter of the string.

Law of lengths – the frequency of a string varies inversely as its length.

Law of tensions – the frequency of a vibrating string varies directly as the square root of the tension.

Latent heat of fusion – the number of calories required to change 1 gram of a substance from a solid to a liquid state without a change in temperature (metric system).

Latent heat of vaporization – the number of calories of heat required to change 1 gram of a liquid into vapor without changing its temperature (metric system).

Longitudinal vibrations – waves in which the particles vibrate along the path which the wave travels.

Loudness – auditory sensation which depends upon intensity, distance, and medium.

Manometer – an instrument used to measure fluid pressure.

Mass – a quantity of matter.

Mass-density – mass per unit volume.

Molecule – the smallest particle of any substance that has the properties of the substance.

Node – region of zero motion.

Noncrystalline – having no definite melting point.

Osmosis – the exchange which takes place between two fluids of different densities when separated by a moist semipermeable membrane.

Overtones – tones produced by a vibrating body having frequencies higher than the lowest or fundamental.

Pascal's principle – the pressure applied anywhere on a confined fluid is transmitted undiminished throughout the body of fluid.

Period – time in seconds required for one cycle.

Pitch – auditory sensation in terms of which sounds may be ordered on a scale.

Pressure – a force per unit area.

Prévost's law – heat lost is equal to heat gained.

Rarefaction – region in a wave motion where the medium is rarefied.

Resonance – increase in the average sound energy because of the reinforcement of sound waves.

Siphon – a tube used to transfer liquids from a higher to a lower level.

Solidification – a change of state from a liquid to a solid.

Specific heat – the number of calories required to raise the temperature of 1 gram of a substance 1 centigrade degree (metric system).

Spherometer – a measuring device used to determine the curvature of a spherical surface.

Standard pressure – 76 cm. of mercury.

Standard temperature – 0°C.

Standing waves – waves apparently standing still due to the interference of two similar waves.

Stratosphere – the layer of the atmosphere immediately above the troposphere.

Sublimation – change from solid to vapor state.

Surface tension – the tendency of a surface to contract.

Sympathetic vibrations – vibrations set up in one object by the vibrations of the same frequency in another object.

Temperature – intensity of heat.

Thermometer – a temperature measuring device.

Train of waves – a series of condensations and rarefactions in a medium.

Transverse vibrations – waves in which the particles vibrate at right angles to the path along which the wave travels.

Troposphere – bottom layer of the atmosphere.

Vaporization – a change from liquid to vapor state.

Velocity – rate of motion in a definite direction.

Velocity ratio – the distance the force moves over the distance the resistance moves.

Venturi tube – a tube with a constriction in which Bernoulli's principle is applied.

Viscosity – friction between layers of a liquid.

Water equivalent – the mass of water which has the same thermal capacity as the body.

Wave length – the distance between two consecutive particles in the same phase.

Weight – the pull of gravity on a mass.

Weight-density – weight per unit volume.

FOR FURTHER READING

Books

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Craff, Graham, Young: *The Elements of Chemistry*. Toronto: Clarke, Irwin Company, 1959.

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May: *There's Adventure in Chemistry*. Chicago: Popular Mechanics.

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Westmeyer: *Successful Devices in Teaching Chemistry*. Portland, Maine: J. Weston Walch.

Periodicals

The Science Teacher, Journal of the National Science Teachers Association, 1201 Sixteenth Street, N.W., Washington 6, D.C.

Science World, 33 West 42nd Street, New York 36, N.Y.

Experiment-of-the-Month Club, P. O. Box 325, Danbury, Conn.

Things of Science, Science Service, 1719 N. Street, N.W., Washington 6, D.C.

Films

Teachers can receive information about new science films by writing to:
The Promotion Division, The National Film Board, P. O. Box 6100, Montreal 3, Quebec.

ACKNOWLEDGMENTS

INDEX

TRANSITION

Period	Group I A		III B	IV B	V B	VI B	VII B	
1								
2	2,1 **3 Li** 6.939	2,2 **4 Be** 9.0122						
3	2,8,1 **11 Na** 22.9898	2,8,2 **12 Mg** 24.312						
4	2,8,8,1 **19 K** 39.102	2,8,8,2 **20 Ca** 40.08	2,8,9,2 **21 Sc** 44.956	2,8,10,2 **22 Ti** 47.90	2,8,11,2 **23 V** 50.942	2,8,13,1 **24 Cr** 51.996	2,8,13,2 **25 Mn** 54.9380	2,8,14,2 **26 Fe** 55.847
5	2,8,18,8,1 **37 Rb** 85.47	2,8,18,8,2 **38 Sr** 87.62	2,8,18,9,2 **39 Y** 88.905	2,8,18,10,2 **40 Zr** 91.22	2,8,18,12,1 **41 Nb** 92.906	2,8,18,13,1 **42 Mo** 95.94	2,8,18,13,2 **43 Tc** (99)	2,8,18,15,1 **44 Ru** 101.07
6	2,8,18,18,8,1 **55 Cs** 132.905	2,8,18,18,8,2 **56 Ba** 137.34	**57-71** See Lanthanide Series	2,8,18,32,10,2 **72 Hf** 178.49	2,8,18,32,11,2 **73 Ta** 180.948	2,8,18,32,12,2 **74 W** 183.85	2,8,18,32,13,2 **75 Re** 186.2	2,8,18,32,14,2 **76 Os** 190.2
7	2,8,18,32,18,8,1 **87 Fr** (223)	2,8,18,32,18,8,2 **88 Ra** (226)	**89-100** See Actinide Series					

NOTE "A" groups consist of elements where the outside ring carries the valence electrons.
"B" groups are the transition groups where a shell other than the outermost is being filled.
Vertical column to the left of the element shows ring structure. The heavy step line represents the division between metals and non-metals. The heavy vertical line separates the inert elements.

Lanthanide Series	2,8,18,18,9,2 **57 La** 138.91	2,8,18,19,9,2 **58 Ce** 140.12	2,8,18,20,9,2 **59 Pr** 140.907	2,8,18,22,8,2 **60 Nd** 144.24	2,8,18,23,8,2 **61 Pm** (147)	2,8,18,24,8,2 **62 Sm** 150.35
Actinide Series	2,8,18,32,18,9,2 **89 Ac** (227)	2,8,18,32,19,9,2 **90 Th** 232.038	2,8,18,32,20,9,2 **91 Pa** (231)	2,8,18,32,21,9,2 **92 U** 238.038	2,8,18,32,22,9,2 **93 Np** (237)	2,8,18,32,23,9,2 **94 Pu** (242)

ELEMENTS SHOWING ELECTRON STRUCTURES

1 **H** 1.0079				2 **He** 4.0026

ELEMENTS

III A IV A V A VI A VII A

| 2,3 5
B
10.811 | 2,4 6
C
12.01115 | 2,5 7
N
14.0067 | 2,6 8
O
15.9994 | 2,7 9
F
18.9984 | 2,8 10
Ne
20.183 |

VIII B

I B II B

| 2,8,3 13
Al
26.9815 | 2,8,4 14
Si
28.086 | 2,8,5 15
P
30.9738 | 2,8,6 16
S
32.064 | 2,8,7 17
Cl
35.453 | 2,8,8 18
Ar
39.948 |

| 2,8,15,2 27
Co
58.9332 | 2,8,16,2 28
Ni
58.71 | 2,8,18,1 29
Cu
63.54 | 2,8,18,2 30
Zn
65.37 | 2,8,18,3 31
Ga
69.72 | 2,8,18,4 32
Ge
72.59 | 2,8,18,5 33
As
74.9216 | 2,8,18,6 34
Se
78.96 | 2,8,18,7 35
Br
79.909 | 2,8,18,8 36
Kr
83.80 |

| 2,8,18,16,1 45
Rh
102.905 | 2,8,18,18 46
Pd
106.4 | 2,8,18 47
Ag | 2,8,18 48
Cd | 2,8,18 49
In | 2,8,18 50
Sn | 2,8,18 51
Sb | 2,8,18 52
Te | 2,8,18,8 53
I
9044 | 2,8,18,18,8 54
Xe
131.30 |

| 2,8,18,32,17 77
Ir
192.2 | 2,8,18,32,17,1 78
Pt
195.09 | | | | | | | 85
At
) | 2,8,18,32,18,8 86
Rn
(222) |

Atomic weight in bracke[t]
denotes the mass numb[er]
of the isotope of longe[st]
known half-life.

| 2,8,18,25,8,2 63
Eu
151.96 | 2,8,18,25,9,2 64
Gd
157.25 |
| 2,8,18,32,24,9,2 95
Am
(243) | 2,8,18,32,25,9,2 96
Cm
(247) |